

## UNIVERSITY OF MADRAS

## B. T. Degree Examination, April 1948

## GENERAL METHODS—(C)

Time : Two and a half hours                      Maximum : 75 marks

*All questions carry equal marks*

1. You are put in charge of a small High School noted for its laxity of discipline. What steps would you take to inculcate a sense of discipline among your pupils?

2. In the New Teaching it is recognised that the pupils ~~should play~~ an active part than in the Old.

Select some topic of study and show how you would try to put into practice the above method.

3. Are children's questions always to be answered?

Discuss with reference to any three types of questions.

4. Illustrate and explain the distinction between Deductive and Inductive types of lessons.

5. Explain, with reference to some topic chosen by you, how you will combine class teaching with group work in a class.

Turn over

# ELEMENTARY SCIENCE

FOR

THE S. S. L. C. COURSE.

II—CHEMISTRY.

BY

Prof. P. LAKSHMI NARASU,

*Retired Professor of Physics.*

THEENATHAYALU & Co.,

PUBLISHERS & BOOKSELLERS,

176, MINT STREET, MADRAS.

*All Rights Reserved.*

*[Rupee One.]*

## Syllabus.

II. *Chemistry*.—Physical and chemical changes mixtures and compounds ; elements and compounds.

Study of the composition of air.

Water—Qualitative and quantitative composition. Impurities in natural waters ; hardness ; pollution of well and river waters and methods of purification of water.

Preparation and properties of oxygen, hydrogen, nitrogen and carbonic acid gas.

The following processes and manipulations to be learnt during the course ; solution, decantation, filtration, precipitation, distillation, evaporation, crystallization.

Non-metals—One method of preparation and the useful properties of chlorine, iodine, sulphur, phosphorus and carbon.

The physical and obvious chemical properties of the following substances to be studied during the course :—

Alum, ammonium chloride, copper sulphate, ferrous sulphate, sulphur, borax, nitre, epsom salt, common salt, potassium chlorate, soda, red oxide of mercury, chalk, quicklime, lime water, phosphorus, potassium, sodium, copper, iron, lead, magnesium, the common acids and alcohol.

---

## Preface.

The method of treatment in this part is the same as in Part I. The pupils are expected to carry out most of the experiments, as only personal observation and handling of things will enable them to understand chemical manipulations and acquire correct ideas of chemical change. Facts and phenomena occupy the foreground; only after the individual facts come the generalisations and laws. As the pupil must be led from the known to the unknown, the known metals are spoken of in a general way before treating of the unknown gases. Which experiments shall be performed by the teacher in the class room, and which in the laboratory by the pupils, have been left to the discretion of the teacher, as a great deal in this matter will depend upon the equipment of the school. The syllabus has not been followed slavishly, but everything has been done to make the book a stepping stone to further study. In conclusion I thank all who have been of help to me in bringing out the book through the press.

P. L. N.

---



# TABLE OF CONTENTS.

---

	PAGE.
I. Single substances and mixtures ...	1
II. Physical and chemical changes ...	12
III. Conservation of mass ...	18
IV. Chemical combination-Compounds ...	20
V. Chemical decomposition-Elements ...	25
VI. Metals ...	29
VII. The composition of the air ...	34
VIII. The composition of the air ...	38
IX. Do. do. ...	43
X. Oxygen ...	48
XI. Water ...	54
XII. The composition of water ...	59
XIII. Hydrogen ...	65
XIV. Reduction. Composition of Water ...	71
XV. Combustion ...	77
XVI. Sulphur-Sulphides ...	81
XVII. Sulphur dioxide-Sulphuric acid ...	84
XVIII. Carbon ...	90
XIX. Carbonic anhydride ...	96
XX. Carbon dioxide in the air ...	99
XXI. Chalk. Potash. Soda. Carbonates ...	103
XXII. Nitrogen-Ammonia ...	105
XXIII. Common salt. Sodium ...	111
XXIV. Hydrogen chloride ...	113
XXV. Chlorine ...	118
XXVI. Iodine ...	122
XXVII. Nitre ...	124
XXVIII. Nitric acid. Nitrates ...	126
XXIX. Bone ash. Phosphorus ...	129
XXX. Acids. Salts. Bases ...	133
XXXI. Mercury. Lead. Tin ...	138
XXXII. Iron. Copper. Zinc ...	141
XXXIII. Aluminium ...	145
XXXIV. Silver. Gold ...	148
Questions ...	150

---

## SINGLE SUBSTANCES AND MIXTURES.

1. **Substance.**—In the world around us are innumerable substances with different properties. Many of these stuffs, such as gold, diamond, coal, cotton, petroleum, &c, are known as **natural** or **raw products**. Others, such as aluminium, paper, charcoal, glass, &c., are prepared by man from naturally occurring substances, and are called **artificial** or **manufactured** products.

Some substances, such as cotton, sugar, butter, starch, are produced solely by the activity of special organs in plants or animals. Such substances are said to be **organic**. On the other hand, those substances that have not their origin in the activities of plants or animals are said to be **inorganic** or **mineral**.

Whether a substance occurs in nature, or is prepared by man, so long as the essential properties are the same we recognise the stuff as the same. For example, sulphur occurs in nature and is extracted by man from other substances. As the properties of both are the same, we treat both as the same stuff.

2. **Single substance and mixture.**—Whether a body consists of one substance, or is made up of different substances, we can know only by an examination of properties.

**Experiment.**—Examine a piece of rock-crystal and a piece of granite. The former is homogeneous and you recognise in it only one kind of stuff with the naked eye and with a magnifying glass. On the other hand, the granite appears heterogeneous even to the naked eye, consisting of white, black and shining parts. We say that rock crystal consists of only one stuff and that granite is a mixture of three different substances.

We have recourse to different methods to decide whether a given material consists of a single stuff or is a mixture of several substances.

**Exp.**—Make an intimate mixture of iron powder (*ferrum limatum*) and flowers of sulphur by grinding them together in a mortar with a pestle (Fig. 1).

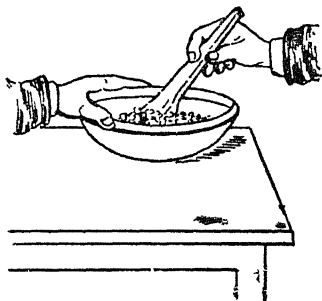


Fig. 1.

Iron powder is grey and the flowers of sulphur are yellow. The mixture appears yellowish grey. Slowly drop a third of the mixture into water. The heavier particles of iron sink faster and collect as a sediment at the bottom, while the lighter sulphur particles lag behind and float as a scum at the top. Thus the two ingredients of the mixture have separated to some extent by **elutriation** or washing. Gold-bearing quartz is crushed to powder and washed with water to separate the heavier grains of gold from the lighter grains of quartz admixed with it.

Blow gently over a second part of the mixture ; observe that the lighter particles of sulphur are blown away from the heavier particles of iron. This process of **winnowing** is, as you know, very often employed to separate the lighter husk from the heavier grain by the blowing wind.

Bring the pole of a magnet close to some iron powder. Observe that the iron powder adheres in a tuft to the magnet. Wipe off the iron powder from the magnet with a paper or cloth, and plunge the end of the magnet in flowers of sulphur. The flowers of sulphur do not adhere to the magnet. A magnet attracts iron, but has no action on sulphur. Now introduce the magnet into the remainder of the mixture of iron powder and flowers of sulphur. The iron adheres

to the magnet, while the sulphur particles fall off as the magnet is lifted. The powder left behind is more yellow than the original mixture. Why? This method of **magnetic separation** is often employed to separate magnetic substances (iron ores) from non-magnetic substances (zinc or copper ores).

**3 Solution—Exp**—Introduce into water some nitre or saltpetre, a substance largely employed in making fireworks and gun-powder. Observe that the substance soon disappears. Taste the water. The brackish taste of the water indicates the presence of the nitre in it. The solid substance has become intimately mixed with the liquid and assumed the liquid form. In this case the solid is said to be **dissolved** in the liquid. The original liquid is called the **solvent**, the dissolved substance the **solute**, and the resulting liquid a **solution** of the substance. Also the process of being dissolved is called **solution**.

A solute may be a solid, a liquid, or a gas, but the solvent is always a liquid. Instead of water, alcohol and carbon disulphide are often used as solvents.

Substances which easily dissolve in water are said to be **soluble**, while those that do not dissolve in water are said to be **insoluble**.

Find which of the following substances are soluble in water—Alum, borax, blue vitriol, chalk, charcoal, copperas, jaggery, nitre, plaster of Paris, quartz, sand.

**Exp**—Put a quantity of nitre in some water in a boiling flask, and stir it up. Soon a point is reached when the liquid will not dissolve any more of the solid. The solution is then said to be **saturated**. Heat the flask with its contents over a flame. Observe that more of the solid dissolves. Generally a **solvent** dissolves more of the **solute**, as the temperature is raised.

A solution which is not saturated is said to be **unsaturated**. If an unsaturated solution contains a proportion of the solute very near that of a saturated solution, it is said to be **concentrated**, otherwise it is said to be **dilute**.

The number of grams of a solute which dissolves in a definite mass, namely, 100 grams of the solvent, to form a saturated solution at a given temperature, is called the **solubility** of the solute in the solvent at that temperature. The solubility of a solute in a solvent depends on the nature of both and on the temperature. Some substances, like potassium permanganate, are very soluble in water, while others, like plaster of Paris, are difficultly soluble. In most cases solubility increases with rise of temperature. The effect of temperature is very marked in the solubility of nitre in water, while the solubility of common salt is nearly the same at all temperatures, and that of slaked lime is less in hot water than in cold. The solubility of nitre at 21°C. is 21.2 and at 100°C., 247 ; the solubilities of common salt for these temperatures are respectively 36 and 39.6.

Some substances are so readily soluble in water that they absorb moisture from the air, and become moist, and even dissolve in the water they have absorbed. Such are caustic soda, caustic potash, and calcium chloride. This property is called **deliquescence**, and the substances are said to **deliquesce**, to be **deliquescent** or **hygroscopic**.

Suppose we have a mixture of common salt and sand, and we desire to separate them from each other. All that we have to do is to treat the mixture with water, until no more will dissolve. Then the soluble part goes into solution leaving the insoluble part behind in suspension. How may we separate the solution from the undissolved part ? This may be done in two ways. On allowing the liquid to stand the insoluble part settles down to the bottom. Then the clear liquid may be **decanted** into another vessel.

4. **Filtration**.—More often solid particles that are suspended in a liquid are removed from the liquid by **filtration**. A filter is a porous body, whose pores are sufficiently large for liquids to pass through, but too narrow for the passage of solid particles. In the household melted ghee, honey and other liquids are filtered

through muslin. In large cities water is filtered through beds of sand.

In chemical experiments filter paper, a pure un-sized paper, resembling white blotting paper, is used. A circular piece of filter paper is folded into half and again into a quarter, and opened out so as to form a cone, and this cone is fitted into a glass funnel (Fig. 2), taking care that the paper never extends above the

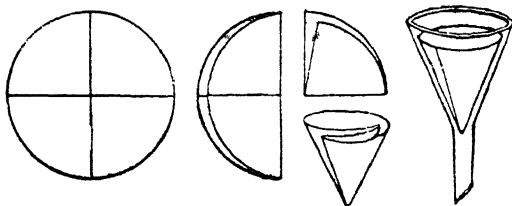


Fig 2.

edges of the funnel, and that the apex of the cone penetrates slightly into the stem. The funnel is placed on a tripod, or on a ring of a retort or funnel stand, or in the mouth of a flask or bottle. The liquid with the suspended particles is poured down a glass rod to the filter paper (Fig. 3). A clear liquid runs through, leaving a residue on the filter. The clear liquid is called the **filtrate**.

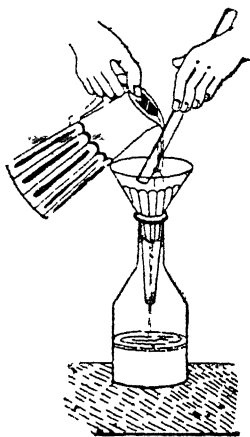


Fig. 3.

Neither by decantation nor by filtration can we completely separate the soluble substance (common salt) from the insoluble sand. This can be effected only by washing. For this purpose pure water (distilled water) is poured on the residue in the filter, till the filtrate no longer contains any traces of the soluble substance.

In washing a chemist employs

on wash-bottle (Fig. 4.) which is filled with distilled water. By blowing into the short tube which points obliquely

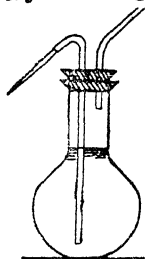


Fig. 4

upwards we obtain a fine stream through the jet, which may be turned in any desired direction by the rubber joint, or by inverting the flask we can pour a larger quantity through the shorter tube.

How can we obtain the dissolved substance (common salt) from the filtrate? This may be done by evaporation, distillation, or crystallisation.

5. **Evaporation.**—**Exp.**—Pour the filtrate into a porcelain dish, and heat the dish on wire gauze over a flame (Fig. 5).

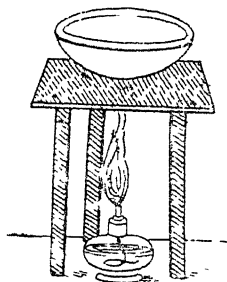


Fig. 5.

Observe that after some time all the liquid has disappeared, leaving a white solid behind as residue. The white solid is the dissolved common salt. The solvent, water, has become changed into vapour. The process by which a solid or liquid changes into vapour is called **evaporation**. Evaporation may occur slowly from the surface, or rapidly throughout the mass with bubbles of vapour rising through the liquid.

In the latter case the liquid is said to boil and the process is called **ebullition** or **boiling**.

**Exp.**—Make a saturated solution of common salt by adding to 50 c.c. of cold water enough salt as to leave after thorough shaking a portion undissolved. Weigh a porcelain dish ( $W$  grams). Pour a portion of the clear liquid into the dish and weigh again. Let the weight be  $W_1$  grams. Therefore the weight of the solution taken is  $(W_1 - W)$  grams. Now evaporate the solution and weigh the dish with the solid residue. Let the present weight be  $W_2$  grams. Then the weight of the dissolved salt is  $(W_2 - W)$  grams. Therefore the weight of water in the solution taken is  $(W_1 - W) - (W_2 - W)$  or

$W_1 - W_2$  grams. Therefore the solubility is  $100 \times (W_2 - W) \div (W_1 - W_2)$ .

For example, 20 c. c. of a solution weighed 24.72 grams and on evaporation yielded 6.5 grams. Therefore the weight of the solvent is 18.22. Therefore the solubility is  $6.5 \times 100 \div 18.22$  or 35.67

The vapour produced by the evaporation of a liquid may be condensed back to the liquid state by the process of **distillation**(=evaporation + condensation).

**Exp.**—Instead of heating the filtrate in an open dish as in the last experiment, heat it in a retort, inserting the neck of the retort in a glass flask (receiver), which is cooled on the outside by pouring cold water (Fig. 6). The invisible vapour, steam, produced by the evaporation of the filtrate, condenses as water, the **distillate**.

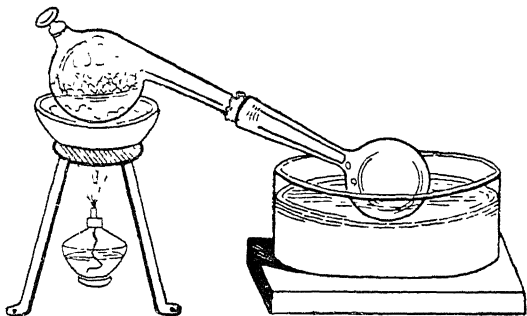


Fig. 6.

When most of the liquid has evaporated, remove the flask, and evaporate the residue to dryness. By separately evaporating some of the distillate in a clean porcelain dish, it can be seen that no dissolved substance is present in the distillate. The distillate is the pure solvent

The process of distillation is largely employed for preparing pure water, for separating more volatile substances (lower boiling point) from the less volatile substances (higher boiling point) mixed with them. For



example, petrol, kerosin oil, paraffin, &c., are separated from petroleum, which occurs in nature, by a series of distillations. So also spirit of wine is prepared from fermented liquor; alcohol is separated from water.

One of the characteristic properties of water is its great solvent action on a very large number of substances. Owing to its great dissolving action it is very difficult to find in nature perfectly pure water. Even the clearest water found on the surface of the earth leaves a residue on evaporation. Rain water, collected in the open, is the purest form of natural water, but it contains some of the gases of the atmosphere in solution, though it does not leave a residue on evaporation. We obtain pure water by distillation.

**6. Crystallisation.**—When a saturated solution of a solid is cooled or evaporated slowly, the dissolved solid separates from the solvent in the form of **crystals**, that is, solids having a definite form bounded by plane faces. The form of the crystals of a substance is a characteristic of the substance, and serves often to identify it. For example, common salt crystallises in cubes, alum in octohedra, nitre in needle shaped prisms, and so forth. The formation of crystals is called **crystallisation**.

**Exp.**—Shake up water in a flask with finely powdered blue vitriol (copper sulphate), until no more goes into solution. We have then a saturated solution of blue-stone. By boiling we can make more of the solid dissolve. On cooling the solution the dissolved substance is deposited as crystals. Filter off the crystals, dry them between pieces of filter paper and observe their plane faces and well-defined angles. Note that each crystal, no matter what its size may be, has the same form.

How is sugar-candy formed from sugar?

Many crystals owe their crystalline form to the presence of water in them, and on heating lose their water of crystallisation and fall to powder.

**Exp.**—Powder some of the crystals obtained in the last experiment and heat the powder in a

crucible over a flame for some time. Observe that the blue powder gradually becomes white. Add a drop of water to some of the white powder. Observe that the blue colour is restored.

**Exp.**—Place some crystals of washing soda in an evaporating dish and cover the dish with an inverted funnel and heat gently on a sand bath (a tin dish containing sand). Observe the drops of moisture condensing on the interior of the funnel and the vapour escaping through the stem (Fig. 7.)



Fig. 7

Substances which have lost their water of crystallisation are said to be **dehydrated**. Some crystalline substances, as washing soda, lose their water of crystallisation on simple exposure to air. They are said to **effloresce**, and the process is called **efflorescence**. By efflorescence clear and colourless crystals of washing soda become powdery and white. Some substances, such as common salt, nitre sugar, which crystallise from their aqueous solutions, do so without any water of

crystallisation. They are said to be **anhydrous**. When some anhydrous crystals, such as common salt, are heated, they crackle or produce sudden sharp splitting sounds. This crackling is spoken of as **decrepitation**. This is due to the bursting of the crystals owing to the evaporation of moisture enclosed in them.

Since all solids are not soluble to the same degree in water, the ingredients of a mixture of solids soluble in water may be separated by crystallisation. When a solution containing a number of dissolved solids is gradually evaporated, the least soluble of the solids separates first and the most soluble last. For example, sea water contains besides common salt small quantities of other solids. Of these the least soluble is

calcium sulphate (gypsum), and this appears first on evaporating the water; the common salt crystallizes out next; and the "mother liquor", that is, the liquid left after the crystallisation of common salt, contains salts of magnesium which are very soluble.

**Exp.**—Mix together one part by weight of blue vitriol with about 6 parts by weight of sodium sulphate. Add to the mixture boiling water so as to produce a saturated solution. Filter the hot solution and allow the liquid to cool. Observe that colourless crystals are first formed. Filter off the crystals wash them in a fine stream of water from a wash-bottle, and dry them between pieces of filter paper. These colourless crystals are crystals of sodium sulphate. The blue colour of the filtrate indicates the presence of the blue-stone dissolved in it.

The method of separation by crystallisation is never perfect, as each crop of crystals will generally contain some admixture of the ingredients which separate before and after it.

Crystals may also be obtained from solutions of substances in solvents other than water. Alcohol and carbon disulphide are also largely employed as solvents. Crystals of sulphur are obtained from its solution in carbon disulphide.

**7. Fusion.**—Fusion or melting is the conversion of a solid into the liquid state by the application of heat. This process may sometimes be employed to separate the ingredients of a mixture.

**Exp.**—Make a mixture of white sand and fine scrapings from a stearine candle. Sand is an infusible substance, and stearine melts easily. Heat the mixture, till the stearine melts. Now press the mixture through a piece of muslin. The melted stearine is pressed out, while the infusible sand remains on the cloth.

By fusion we may separate two substances, such as sulphur and bee's wax, which melt at different temperatures. The mixture is heated to a temperature

higher than the melting point of one of the substances but lower than that of the other.

Just as crystals grow from a cooling or evaporating solution, so also they can grow from a molten mass which is slowly solidifying.

**Exp.**—Half fill a fire-clay crucible with powdered sulphur, and heat it slowly on a tripod, till the sulphur melts.

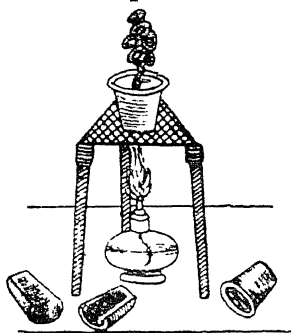


Fig. 8

Allow the crucible to cool on the tripod, taking away the flame (Fig. 8). As soon as a crust forms make a hole in it by means of a glass rod, and holding the crucible with tongs, pour out the remaining liquid through the hole. When the crucible is cold, break away the crust and observe the needle-shaped amber coloured crystals.

**8. Sublimation.**—Sometimes a volatile solid, such as sal-ammoniac, gets mixed with a refractory solid like sand, which is not volatile. In such a case we may resort to the process of **sublimation** for separating the one from the other. In sublimation a solid is directly converted into vapour and is again condensed to the solid state without passing through the intermediate liquid state.

**Exp.**—Heat some camphor in a retort on a sand bath, the camphor soon disappears. At the same time the colorless vapour of camphor condenses in the colder parts of the retort in the form of fine crystals. The camphor has **sublimed** and the crystals form the **sublimate**.

**Exp.**—Gently warm a mixture of sand and sal-ammoniac in a test tube. The sal-ammoniac is converted into vapour and the vapour solidifies in contact with the colder parts of the tube

## II

### PHYSICAL AND CHEMICAL CHANGES.

In nature, in the house-hold, in the manufactory substances are made to go through various processes, which produce changes in their properties. Some of these changes form the subject of study in physics, while others are dealt with in chemistry. To understand clearly the difference between these two classes of change, let us perform a number of parallel experiments.

**Exp.**—You are acquainted with sulphur. It is a tasteless, odourless, brittle, yellow solid, insoluble in water. Heat a piece of sulphur in a test tube. As the temperature is raised, the sulphur changes in form. It melts at  $110^{\circ}$  C. and forms a thin amber-coloured liquid. Pour a portion of the liquid into cold water, and observe that it solidifies into a hard solid. Continue to heat the remainder of the liquid. It grows darker, becomes black and so viscid that it cannot be poured out. Heat it still further. The viscid liquid becomes mobile and lighter in colour. Continue the heating, till the liquid boils. Observe the orange-red colour of the vapour of sulphur, which sublims on the cold sides of the tube to a yellow powder, known as **flowers of sulphur**. By heating and cooling the sulphur has passed through changes of state. We still recognise it as the same stuff.

Once more heat a piece of sulphur but this time in an open dish. The sulphur melts as before. But after a time it catches fire and burns with a blue flame producing a colourless gas of a suffocating odour, which

you recognise as the smell of burning sulphur. When the sulphur has completely burnt away, no residue is left. The gas evolved does not deposit the original sulphur. We no longer recognise in the gas produced by the burning the original stuff, namely, sulphur. The sulphur has disappeared, and lost its identity. It has not merely acquired a new property, but it has become so completely changed that the substance produced is at once recognised as substantially different from sulphur.

**Exp.**—Dissolve a little sugar in a test tube one-fourth full of water. Dip a glass rod into the liquid and taste it. Observe that the character of sugar, namely, sweetness, is in the water, which by itself is tasteless. Dip the rod into the liquid again, and hold it over a flame. As the water evaporates, a white solid appears. Taste it. You still recognise the white solid by its sweetness to be nothing else than the sugar originally dissolved in the water. By its solution in water sugar has suffered a change of form but not of substance.

Heat some of the same sugar in a test tube. It melts forming a pale yellow liquid. The liquid rapidly darkens, and evolves a vapour which escapes out of the tube. On applying a flame to the vapour it burns with a bright flame at the mouth of the tube. But the greater part of it condenses into drops of water on the cold sides of a beaker held over the mouth of the tube. The residue in the tube is found to be a black mass which does not undergo any further change, however long the tube may be heated. This black mass has no sweet taste, does not dissolve in water, nor does it produce sugar. It is nothing but charcoal. By heating sugar has undergone a substantial change and become broken up into other substances, such as water and charcoal.

The decomposition of a substance by heat and the distilling off of the volatile products thereby produced is called **destructive distillation**; sugar in the above experiment has suffered destructive distillation.

**Exp.**—Bring a rod of wrought iron near some iron tacks. Observe that there is no action. Coil round the rod some insulated copper wire and connect the ends of the wire with the poles of a voltaic battery. Now bring the tacks near one end of the rod. Observe that the tacks adhere to the end in a tuft. Under the influence of the electric current flowing through the coil of copper wire the rod has become magnetised, that is, it has acquired the property of attracting iron. Disconnect the ends of the coil from the poles of the battery. Observe that the tacks fall off. The rod has lost the newly acquired property. The change is only temporary. We recognise that the rod still consists of the same stuff iron.

Brighten one end of the rod by rubbing it with sand paper. Heat the bright end to redness and plunge it into water. Observe that the surface has lost its brightness. Repeat the process a number of times. Hammer the iron and note the reddish brown scales falling off from its surface. While metallic iron is hard and readily attracted by a magnet, the reddish scales are brittle and not attracted by a magnet. The scales resemble the rust formed when iron is exposed to moist air. The scales form a substance entirely different from the iron.

**Exp.**—You know that when a candle burns, it completely disappears. What is the change it has undergone? Is it a physical change, or a chemical change?

In a candle the burning material is either bee's wax, stearine, tallow, or paraffin. When the wick is lighted, the burning of the wick produces around it a cup-like depression which is filled with the melted material. The molten liquid can be poured out and on cooling it solidifies. You recognise the solid at once to be no other stuff than the material of the candle. The melting has therefore produced only a temporary change but the substance still remains the same.

Burn a candle and hold over it an inverted dry tumbler (Fig. 9). Observe that the tumbler becomes dim;

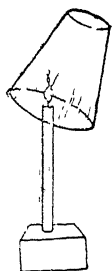


Fig. 9.

as does a glass plate when you breathe on it. You know that when you breathe on a glass plate, the moisture from your breath condenses on the plate and causes its dimness. Similarly the dimness of the tumbler must be due to the moisture produced by the burning of the candle.

Fix a small candle in the cup of a deflagrating spoon, light it and introduce it into a large flask (Fig. 10). Observe that the candle burns for some time and then goes out. When the candle has gone out, withdraw it carefully from the flask, light it again, and reintroduce it into the flask. It goes out at once, showing that the gas now present in the flask is entirely unlike the ordinary air that was in the flask before the burning, since it does not allow a candle to burn in it.

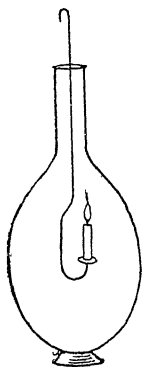


Fig. 10.

Pour some lime water (made by shaking slaked lime with water and decanting the clear liquid after allowing the undissolved lime to settle down) into a flask containing ordinary air and shake it. You observe nothing special. The lime water remains clear. Now pour some lime water into the flask in which you have burnt the candle and shake it. The lime water has turned milky. Now lime water is a **reagent** for carbonic acid gas and the milkiness or turbidity of lime water is a **reaction** of that gas. Hence the milkiness of the lime water poured into the flask indicates the presence in the flask of carbonic acid gas due to the burning candle.

Thus in the burning of a candle nothing has disappeared, but the solid stuff forming the candle has changed into two completely new substances, water and carbonic acid gas.

From these four pairs of examples we learn that substances can undergo two kinds of change :



1. Those which confer on a substance new properties without altering those properties that make it a particular stuff. Such changes are called **physical changes**. A physical change is often only temporary, and although for a time the original properties may disappear, they reappear after the removal of the cause producing the change, so that the substance is recognised as the same after the change as before it.

2. Those which alter completely the properties of a substance so that what results from the change is found to be a new stuff entirely different from the original stuff. Such changes are called **chemical changes**. A chemical change is a permanent change, which affects the essential nature of a substance. The removal of the cause of a chemical change does not restore the original properties of the substance. The formation of one or more new substances differing essentially from the original substance is the special characteristic of a chemical change.

Explain which of the following changes are physical and which chemical: The burning of wood, the solution of common salt in water, the melting of ghee, the souring of milk, the putrefaction of eggs, the formation of a rainbow in the sky, the fermentation of molasses, the formation of steam from water, the tarnishing of tinsel, the fall of rain, the rusting of iron, the ignition of a match.

A chemical change often involves the action of two or more substances on one another.

**Exp.**—Put some copper (say, a pice) in a porcelain dish and pour some nitric acid on it. The copper disappears with the evolution of a brown gas, leaving behind in the dish a green solution. The acid dissolves the copper, but it is not a case of true solution, like the solution of common salt in water. By the action of the acid on the copper a chemical change has taken place and one of the new substances produced has dissolved in the acid. Such a solution may

be treated as **chemical solution**, though it is usual to speak of it also as solution.

The substances, copper and nitric acid, which have acted on each other, are said to **interact** or **react** on each other, and the change taking place is spoken of as a **reaction**. The name reaction is often reserved only for such chemical changes as are indicative of special substances. For example, the formation of a blue solution by the action of nitric acid on a metal indicates the presence of copper, and is spoken of as a reaction of copper. The milkiness of lime water is a reaction for carbonic acid gas.

The different changes which substances undergo furnish a means of classifying the properties of bodies into physical and chemical. Such properties as accompany physical changes are called **physical properties**, while those that require a chemical change for their manifestation are called **chemical properties**. For example, its yellow colour, its crystalline form, its brittleness, its density, its solubility in carbon disulphide are the physical properties of sulphur, while its combustibility is a chemical property.

State some of the physical and chemical properties of the following : iron, paper, sugar, wood, magnesium, sulphur, copper.

Physical properties and physical changes form the subject-matter of **physics**, while chemical properties and chemical changes are studied in **chemistry**. As physical changes often accompany chemical changes, and as substances are recognised by their physical as well as chemical properties, a knowledge of physics is indispensable for a thorough understanding of chemistry. Physics and chemistry are intimately connected and this connection becomes stronger as they advance.

---

### III CONSERVATION OF MASS.

**Exp.**—On one scale pan of a balance place some loaf sugar on a piece of paper and a beaker half full of water, and counterpoise them together by weights in the other scale pan. Now drop the sugar into the water placing the paper back on the pan. After the sugar has dissolved in the water reweigh. Observe that no change of mass has taken place, though the sugar has disappeared in the water. In the physical change that has taken place there is no change of mass. Remember that mass is proportional to weight.

**Exp.**—Cut a thin piece of phosphorus and put it into a test tube partly filled with water. Carefully warm the tube and observe that the phosphorus melts long before the water begins to boil. Phosphorus melts at  $44^{\circ}\text{C}$ . to a colourless liquid.

Phosphorus is a yellowish substance sold generally in the form of sticks of the thickness of a finger. On account of the readiness with which it catches fire in air it has to be always kept and handled under water. It must be taken out of its bottle with forceps and cut under water into bits. A bit should be lifted out of the water either with forceps or a knife—never with the hand—and placed on filter paper and carefully dried by gently pressing—not rubbing—the filter paper on it. Phosphorus should be handled with care, as a small piece of burning phosphorus, coming in contact with the skin, burrows deep into the skin and flesh and produces burns which do not rapidly heal.

Place a dry bit of phosphorus on a tin-plate and touch it with a hot iron wire. Observe that it is immediately ignited and burns with a dazzling white light, evolving dense white fumes. Even at ordinary temperatures phosphorus gives off in contact with the air a white fume with a garlic odour. Is the burning of phosphorus a physical change or a chemical change?

In a round bottomed flask spread a thin layer of dry sand and lay on the sand a dry bit of phosphorus.

Close the flask with a rubber stopper. Then carefully weigh the flask with its contents. Now warm the flask. The phosphorus catches fire and burns. The phosphorus soon disappears leaving a white smoke behind. The phosphorus has undergone a chemical change. Allow the flask to cool and reweigh it. Observe that there is no change of mass (weight).

**Exp.**—Into a thin glass beaker (100 c.c.) pour 50 c.c. of a solution of lead nitrate, and into another of the same capacity pour 25 c.c. of a solution of potassium iodide. Place the beakers with their contents on the left pan of a balance and counterpoise them. Now pour the liquid in one beaker into the other. Observe the separation of a bright yellow solid from the mixture of the two liquids owing to the chemical action between them. Replace the beakers on the pan and reweigh. Observe that there is neither an increase nor a diminution of mass.

The separation of an insoluble solid, as in the last experiment, by a chemical reaction between two solutions, or between a solution and a gas, is called **precipitation**, and the solid that is formed is called a **precipitate**. The yellow solid is a precipitate of lead iodide.

All the above experiments illustrate an important characteristic of all changes, whether physical or chemical. They prove that the same mass persists after a change as before the change. This characteristic is summed up in the **law of the conservation of mass**, which may be stated as follows :

**The total mass (or weight) of the substances taking part in a chemical reaction is equal to the total mass of the new substances formed by it.**

This law looks very simple and plain with the aid of the balance. But, before the time of Lavoisier, it did not meet with general acceptance. It was believed that when substances burned the stuff in them ceased to exist. For example, when a candle burned, it was thought that the material of the candle was utterly lost.

---

## IV

### CHEMICAL COMBINATION-COMPOUNDS.

**Exp.**—Intimately mix 4 grams of flowers of sulphur with 7 grams of iron powder (*ferrum limatum*) by grinding the two together in a mortar. Observe that the resulting powder has a greyish green colour. This powder is called a **mechanical mixture** or simply a **mixture**.

Examine a portion of the mixture with a magnifying glass. The yellow particles of sulphur are easily distinguishable from the greyish particles of iron.

Pass the pole of a magnet through the mixture. The iron easily adheres to the pole leaving the yellow particles of sulphur behind. Thus the iron is easily separated from the sulphur.

The particles of iron may be separated from those of sulphur also by washing, or winnowing, or better still by shaking up the mixture with carbon disulphide. Carbon disulphide is a colorless liquid with an unpleasant odour, highly inflammable and volatile. Be careful that there is no flame in the neighbourhood when dealing with carbon disulphide, as the vapour given off by the liquid readily catches fire.

Though sulphur is insoluble in water, it dissolves readily in carbon disulphide. By shaking the mixture of iron powder and flowers of sulphur with carbon disulphide, the sulphur dissolves in the carbon disulphide, leaving the iron unchanged. By filtration the iron is separated from the solution. On evaporating the solution we get back the sulphur in the form of crystals. Note that the shape of these crystals is different from that of the crystals obtained by fusion.

Treat some of the mixture with dilute sulphuric acid; the iron goes into solution, leaving the sulphur behind. At the same time a gas with a slight odour is

produced. Filter the solution and observe its green colour. Boil the filtrate and allow it to cool. After sometime green crystals are deposited. These are crystals of **green vitriol**. Green vitriol is used largely in ink making and dyeing.

Put the remainder of the mixture in a test-tube. Support the test-tube in a clamp and heat its lower end, till it begins to glow there. Now remove the flame. The glow becomes more intense and spreads slowly through the whole mass. The heat generated in this change often melts the glass tube. After the tube has cooled, break it and observe the black solid left behind. Powder this solid, and examine it under a magnifying glass. All the particles are of the same colour and appearance. No yellow particles can be distinguished from black particles, as in the case of the mixture.

Any process employed for the identification of a substance is called a **test** for that substance. Now apply various tests to detect the presence of iron and sulphur. See if the iron can be separated from the black powder by passing the pole of a magnet through it, or if sulphur can be dissolved out by means of carbon disulphide. The mechanical processes, effective in the case of the mixture, are found to be not serviceable in the case of the substance produced by heating the mixture.

Add a little dilute sulphuric acid to the black powder. Observe that it dissolves almost completely, generating at the same time a foul-smelling gas.

Hence the substance produced by heating a mixture of iron powder and sulphur is a new substance entirely different in properties from both iron and sulphur.

When a single substance with new properties is produced by the union of two or more different substances, the process is called **chemical combination**, and the new substance produced is called a **chemical compound** and is said to be composed or compounded of the original

substances, by whose union it is produced. The latter are spoken as the **components** of the compound.

By the chemical combination of iron and sulphur we obtain the black body as their compound. Iron and sulphur are the components of the black compound (**iron sulphide**).

A substance can undergo chemical combination only by its union with one or more other substances. Hence when a definite mass (weight) of a given substance undergoes a chemical change, and there is an increase of mass, the inference is that chemical combination has taken place. The increase of mass indicates that some new substance has been added to it.

Chemical combination is generally accompanied by the **generation of heat**.

**Exp.**—Put in a dish some lumps of fresh quicklime and over it pour some water. The quicklime falls to powder being converted into slaked lime with evolution of a large quantity of steam. Water combines with quicklime forming slaked lime. The heat evolved in the slaking of lime makes the water boil and produces steam.

Quicklime, as you know, is made from limestone, marble, or shells, by strongly heating them in a kiln. Slaked lime dissolves to a small extent in water. Its solution, **lime water**, is used, as you have seen, as a test for carbonic acid gas. For building purposes we use **mortar**, which is a mixture of slaked lime, sand and water. **Whitewash** is prepared by mixing plenty of slaked lime with water.

Sometimes the heat evolved in combination is so great as to make the products glow. We then have light, as in the combination of iron with sulphur.

An important characteristic which differentiates a compound from a mixture is that a **compound contains the components of which it is composed in a fixed proportion**. For example, the compound of iron and sulphur formed by heating a mixture of flowers of sulphur and iron pow-

der always contains four grams of sulphur for seven grams of iron.

**Exp.**—Make a mixture of flowers of sulphur and iron powder in the proportion of 4 grams of the former to 7 grams of the latter. Add to it an excess of sulphur or iron, and heat the mixture. Observe that the glow is less intense, or does not occur at all. Heat the mixture for a sufficiently long time, and observe that from the resulting stuff the substance added in excess is separated by the same means as from a mechanical mixture.

We may summarise the differences between a mechanical mixture, a solution, and a chemical compound as follows:

A **mechanical mixture** can be separated into its constituents by mechanical or physical methods. It exhibits most of the characteristic properties of each constituent. It does not appreciably alter in properties, if the proportion of the constituents is slightly varied.

A **solution** is a homogeneous mixture of two or more substances, of which one is generally a liquid. It can not be separated into its constituents by mechanical means. Some properties of solutions are entirely different from those of its constituents. Its properties vary continuously with the proportions of the constituents between certain limits.

*Alloys* obtained by melting metals with one another, and *amalgams* obtained by dissolving other metals in mercury, belong to the class of solutions.

A **chemical compound** can not be separated into its components by mechanical or physical means. It does not exhibit the characteristic properties of its components. It always contains components in a fixed proportion by mass (weight).

**Questions**—1. Give reasons for regarding gunpowder as a mixture, and not a compound of sulphur, charcoal, and nitre. (Nitric acid can be separated from the other two constituents by throwing the powder in hot water and



filtering. Sulphur can be separated from charcoal by its solution in carbon disulphide).

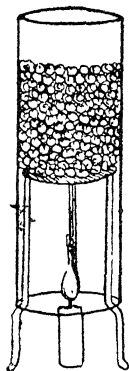


Fig. 11.

2. Take a cylindrical lamp chimney and fit it near its middle with a piece of wire gauze and fill the space above the gauze with soda-lime (a mixture of caustic soda and slaked lime, capable of absorbing water and carbonic acid gas coming into contact with it). Place the lamp chimney over a candle (Fig. 11), and transfer the whole to one pan of a balance and counterpoise it. Now light the candle. Why does the pan sink as the candle goes on burning? What is the inference?

---

## CHEMICAL DECOMPOSITION—ELEMENTS

**Exp**—Into a test-tube of hard glass introduce some red ash of mercury, known among druggists as **red precipitate**. Weigh the whole. Heat the ash in a strong flame. Observe that the colder parts of the tube get coated with a grey metallic coating. Introduce into the tube a red hot splinter (Fig. 12). Observe that the

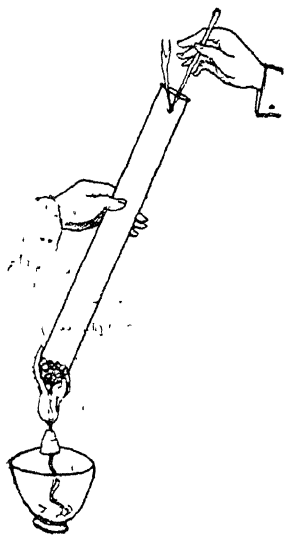


Fig. 12.

splinter is rekindled. **The rekindling of a red hot splinter is a test for a colourless, inodorous gas, called oxygen.** This gas must have come out of the red ash. When the tube has cooled, rub the grey metallic coating on the sides of the tube. Observe that the particles of the grey coating unite into shining drops, which you recognise to be globules of mercury. Re-weigh the tube. There is a diminution of weight. Loss of weight implies the removal of material; it is due to the escape of oxygen from the ash. By the action of heat the red ash of mercury is broken up into metallic mercury and oxygen, two substances entirely

different in properties from the original red ash.

**Exp.**—Examine some crystals of the white salt, sold in the bazaar as **chlorate of potash**. Chlorate of potash is used largely in fireworks and in the preparation of explosive mixtures. A mixture of equal parts of this stuff and sulphur or realgar (an orange coloured compound of sulphur) detonates on hammering,

Chlorate of potash is the most important ingredient in the heads of safety matches. Rub a large crystal of chlorate of potash on the rubbing side of a box of safety matches. Note what happens. The chemical name of chlorate of potash is **potassium chlorate**.

Put some crystals of potassium chlorate in pure water. Observe that the crystals slowly dissolve in the water. When they have dissolved, add to the solution a few drops of a solution of silver nitrate (*lunar caustic*). Observe that there is no apparent change. Place some dry potassium chlorate in a flask and weigh the whole. Now heat the potassium chlorate until it melts. Observe that bubbles of gas are given off by the liquid. Introduce a red hot splinter into the flask, taking care not to touch the liquid. The red hot splinter bursts into flame, showing that the gas evolved by the molten chlorate is oxygen. Continue the heating as long as any gas is produced. Now allow the flask to cool and reweigh. There is a diminution of mass (weight) owing to the gas that has been given out. Now add water to the white solid residue in the flask. Note that it easily dissolves. To the solution add a few drops of a solution of silver nitrate. A white insoluble precipitate is formed. **The solution of silver nitrate is a reagent for a chloride in solution, and the presence of the chloride is indicated by the formation of the white precipitate.** The white residue produced by strongly heating potassium chlorate is therefore a chloride, differing in properties from the original chlorate. Thus by heating potassium chlorate it is broken up into two new substances, namely, a colourless, odourless gas, oxygen, and a white solid, potassium chloride.

**Exp.**—Test some powdered chalk (or limestone) with moist red litmus paper. Nothing happens. Powder the chalk and put some of the powder in drinking water. Filter and taste the filtrate; it is almost tasteless. Place some powdered chalk on a platinum foil (or asbestos paper) and weigh the whole. Now heat the foil in a hot flame till the powder becomes red

hot. Allow it to cool and reweigh. Observe the loss of weight indicating that some thing has escaped from the chalk on heating. Now put some of the residue on moist red litmus paper. The paper turns blue, where it has been touched by the residue. **The blueing of red litmus is a test for alkaline substances and is called an alkaline reaction.** Hence the residue obtained by strongly heating chalk is alkaline.

Add the remainder of the residue to some drinking water and shake well. After allowing the powder to settle, taste some drops of the clear liquid. You notice the taste of lime water. In heating chalk you have effected the same change as is effected in a lime kiln. The chalk has become broken up into a colourless gas and quicklime.

In all the above examples a chemical change has taken place. A single substance by itself gives rise to two or more new substances. Such a change is called **chemical decomposition.**

Chemical decomposition is the reverse of chemical combination. In combination a compound is produced from its components. On the other hand, in decomposition a compound is broken up, resolved, or decomposed into its components. When a substance enters into combination with another, it suffers an increase of mass (weight) owing to the addition of new material. But when a substance undergoes decomposition, each of its components weighs less than the original substance.

Of the many innumerable single substances known to us many can undergo change by decomposition, that is, can be broken up into new substances. But there is a limited number of single substances, which can undergo chemical change only by combination with others but not by decomposition. These single substances are called **simple substances** or **elements**. They cannot by any known methods be separated into two or more different substances.

We know at present 92 elements. Many of them

are only obtainable in small quantities. They are divided into two classes, namely **metals** and **non-metals**. Metals generally possess what is called **metallic lustre** and are **good conductors of heat and electricity**. Elements like sulphur, phosphorus, oxygen, carbon, which do not possess the above mentioned properties of metals, are known as non-metals.

By the **composition of a compound** is meant the elements forming the compound together with their per-centages by weight.

The composition of a compound may be ascertained by **synthesis**, that is, forming it from its elements, or by **analysis**, that is, resolving it into its elements.

The so-called elements of the ancients—earth, water, air, fire, space—are not elements in the chemical sense. Earth is no definite substance, the materials forming the crust of the earth being different in different localities; water is a chemical compound; air is a mixture of gases; fire consists of glowing solids and gases; space is no substance at all.

Whether a substance is simple or compound cannot be decided without considering a definite mass of the substance and the masses of the products which result from its chemical changes. The products resulting from the chemical changes of a given mass of an element always have a greater mass, that is, always weigh more than the weight used of the element. Before the time of Lavoisier (1777) chemical changes were not investigated by means of the balance. Hence such substances, as metals, which we now regard as elements, were treated as compounds. Metals were supposed to be compounds of the calces or ashes they leave behind on burning and an inflammable something, called *phlogiston*. Phlogiston was supposed to escape from burning bodies. This explanation of burning was proved to be false by Lavoisier, who first investigated the phenomena of ordinary burning by means of the balance and established the law of the conservation of mass.

## VI METALS.

Among the elements the metals form a distinct class. Some of them, being in common use, are well-known to you. It will be useful to call to mind here some of their properties already known to you.

1. **Iron.**—On account of its multifarious uses iron is the most important of metals.

There are three varieties of iron in use, namely, cast-iron, steel, wrought-iron. None of these is pure iron.

*Cast-iron* has a grey colour and a coarsely granular structure; it is brittle and melts at  $1200^{\circ}\text{C}$ .

Parts of machines, fenders, fire-grates, pillars, lamp-posts, brackets, water pipes, gas pipes, bedsteads, cooking vessels, ovens, railings are made of cast-iron.

*Steel* has a bright grey colour and a finely granular structure. It is elastic, malleable, weldable. Through rapid cooling it becomes so hard that it scratches all other metals. It can be permanently magnetised.

Steel is used for making nibs, weapons, guns, armour-plates, tools (files, chisels, saws, hammers), needles, cutting instruments (knives, scissors), screw nails, rails, springs, movable parts of machines.

*Wrought-iron* differs from steel in its fibrous structure; its fracture is therefore not smooth as in steel, but shows small projections, as if it were composed of fine wires bundled together. It is very malleable and can be rolled into plates and drawn into wires.

**Pure iron** does not occur in the earth, except in the meteoric iron that has fallen from the heavens.

The large employment of iron for various purposes depends on its cheapness, fusibility, elasticity and malleability.

2. **Copper.**—Copper has a striking red colour. On account of its malleability it can be beaten into thin leaf (tinsel gold) and drawn into very thin wires. It is a very good conductor of heat and electricity. It melts at  $1100^{\circ}\text{C}$ .

Copper plates and wires are used for various household and technical purposes.

Copper is too soft for some purposes and cannot be used for casting. But many of its alloys are harder and can be cast. The most important alloys of copper are :

*Brass*, an alloy of copper and zinc, possessing a bright yellow colour. It is not so malleable as copper, but harder and more easily fusible. Mention some-things made of brass.

*Bronze*, an alloy of copper and tin. Its colour varies from golden-brown to whitish-yellow according to the proportion of copper. Medals, coins, cannon are made of bronze. Bell-metal is a variety of bronze.

In America pure copper is found in abundance.

3. **Lead.**—Lead is characterised by its high specific gravity (11'4). It has a bluish grey colour, its fresh surface is very bright. It is so soft that it can be scratched with one's nail, and also marks paper. It is very malleable and can be beaten into thin foil. It melts at  $230^{\circ}\text{C}$ .

Lead is largely employed for making gas-pipes, loading articles that stand on feet, covering roofs, joining iron pipes. Though lead is a cheap metal, yet on account of its poisonous character it cannot be used for coating utensils for house-hold purposes.

4. **Zinc.**—Zinc has a bluish-white colour with a lustrous surface when polished, but loses its brightness soon. It melts at  $420^{\circ}\text{C}$ .

It is largely employed as sheets. It is used for coating iron in order to protect the latter from rusting. This is done by plunging iron into molten zinc. The iron is then known as *galvanised iron*.

5. **Tin.**—Tin is a white metal with a bright lustre. It is softer than zinc but harder than lead. It melts at  $230^{\circ}\text{C}$ . On account of its malleability it can be beaten into thin leaf (tin foil). Its silver-white colour, easy fusibility and great malleability render it highly useful. Tin foil is used for folding chocolate, tobacco, eatables. Cooking utensils of copper are coated on the inside with tin to prevent the formation of verdigris. Ordinary looking glasses are made by coating glass plates with slightly amalgamated tin foil. The plumber makes use of an alloy of tin and lead in equal parts for soldering. Ordinary tin ware consists of sheet iron coated with tin. This is effected by plunging clean sheets of iron into molten tin.

6. **Nickel.**—Nickel has a white colour and a lasting lustre. It is employed for making coins. These consist of an alloy of nickel and copper. Copper and iron articles are coated with nickel to give them a fine appearance.

7. **Aluminium.**—Aluminium is silver-white and has a very low specific gravity ( $2\cdot6$ ). It melts at  $700^{\circ}\text{C}$ . Like iron it can be forged, welded and beaten into thin leaf. It is used for making keys, cooking utensils, parts of machinery, and household articles.

8. **Magnesium.**—Magnesium is silver-white. It is sold in the form of wire, ribbon, or powder.

A piece of magnesium wire or ribbon, when one end of it is held in a flame, burns with a dazzling white light. It is therefore used in fireworks. Magnesium powder, when strewn on a flame, burns off with a momentary flash. Hence it is used for flash-light in photographing in darkness.

9. **Mercury.**—Quicksilver or mercury is the only metal which at ordinary temperatures is liquid. It is silver white, has a high lustre, and a high specific gravity ( $13\cdot6$ ). It solidifies at  $-39^{\circ}\text{C}$  and gives off vapour even at ordinary temperatures. Its vapour is highly poisonous.



All metals, with the exception of iron, are soluble in mercury and form with it butter-like substances, called **amalgams**.

Mercury is employed for filling barometers and thermometers. Tin amalgam is used for coating mirrors.

10. **Silver**.—Silver has a pure white colour and a lustre. It melts at  $1000^{\circ}\text{C}$ . It is very ductile and malleable. Its hardness is greater than that of gold. In order to make silver articles sufficiently hard for ordinary use silver is alloyed with copper.

Silver is used for making coins, jewels, table service, cases.

11. **Gold**.—Gold is the most valuable of metals. It possesses a fine yellow colour with a strong lustre. It is highly malleable. It can be beaten into such thin leaf that 10000 of them laid on one another will be 1 mm. thick. Its specific gravity is 19.5. It melts at  $1200^{\circ}\text{C}$ .

It is employed largely for coinage and jewelry. As gold is soft, it is alloyed with copper. Pure gold is said to be 24 carats fine. 18 carat gold has 75 per cent fineness, and so on. The British gold coinage contains 11 parts by weight of gold and 1 of copper.

Brass, copper and silver articles are covered with a layer of gold either by fire-gilding or by electro-plating. In the former process the article is rubbed with gold amalgam and then heated; mercury flies off leaving a layer of gold behind.

12. **Platinum**.—Platinum is rarer than gold but not so costly. It is a silver white metal of high specific gravity (21.5). It melts at  $1770^{\circ}\text{C}$ , a temperature far higher than any coal fire. It is very ductile and can be drawn into wire as thin as a spider thread. It occurs in nature in the pure state in the form of small granules.

The standard meter, the standard kilogram are made of platinum. It being the only metal which can

be fused into glass, it is used in incandescent lamps for conducting the electric current. It is used much in the chemical laboratory in the form of wire, foil and crucible.

All metals have some common characteristics. They exhibit ordinarily, or when polished, a sheen: which is called **metallic lustre**. They are malleable and ductile. They generally conduct heat and electricity well.

— —

## VII

### THE COMPOSITION OF THE AIR.

1. **The Air.**—The atmospheric air, which surrounds us on all sides, is a colourless, odourless and tasteless gas. When it is at rest, we do not perceive its presence. Only when it is in motion, or when our bodies are rapidly moving through it, do we notice its existence. It can be rendered visible by displacing it by water. As it possesses extension and mass (weight), it must be considered a substance. That the air, like every other substance, has mass, is proved by the following experiment:—

**Exp.**—Fill a round bottomed flask nearly to a third of its volume with water, and close it tightly with a rubber stopper through which passes a short glass tube ending in a piece of rubber tubing carrying a pinch-cock. Open the pinch-cock and boil the water for sometime to make sure that the issuing steam has driven out the air. Close the pinch-cock and quickly remove the flame. Allow the flask to cool completely. Now counterpoise it on a balance. Then open the pinch-cock. The air rushes in with a hissing sound and the balance sinks on the side of the flask. The increase of mass is due to the mass of the air which has entered the flask to make up for the air driven out by the steam from the boiling water. Precise experiments show that 1 cubic meter of air (at 0° C. and 76 cm. pressure) has a mass of 1.293 kilograms.

2. **Change of metals in the air.**—(a) Iron, copper, lead, zinc, and tin, which have been long lying in the air, undergo a change in their surface. They lose their original colour and brightness. These changes take place most rapidly in potassium and sodium. Gold, silver, platinum, mercury and a few other metals are unaffected by the air. That is why they are known as the noble metals.

### **Metals tarnish or rust in the air.**

Iron rust is a reddish brown powder. Lead, zinc and tin become covered with a grey coating. Copper slowly darkens, becomes gradually black, and finally green.

(b) If the surface of a polished iron object be covered with a thin layer of oil or fat, the metal does not rust. Tin foil enclosed between two glass plates retains its brightness for a long time. When preserved under petroleum, the freshly cut surfaces of potassium and sodium are silver-white and shining.

### **If air is excluded from metals, they do not rust.**

How are weapons and the parts of machinery protected from rusting? Why are iron doors, posts, beams, etc., lacquered or coated with oil-paints? Tin foil is preserved by laying it between sheets of paper closely laid on one another. Iron wire is coated with slowly rusting zinc. Cooking utensils of iron are tinned (sheet tin), because tin rusts less readily than iron. Fine brass measuring instruments are varnished with transparent lacquer varnish.

### **(c) The rusting takes place much faster, when metals are heated in the air.**

Iron becomes mottled and becomes finally covered with greyish black scales, which easily fall off.

Copper first shows rainbow colours and turns gradually black. If it be quickly plunged in water, the black layer peels off.

Molten lead in air becomes covered with a grey scum, which is yellow when hot. On stirring the molten metal with an iron wire the whole changes into a yellow powder, called **litharge**. On still further heating this yellow powder at the melting point of lead, it changes in colour, becoming a red powder, called **red lead**.

If a piece of sodium be heated in an iron spoon, it melts and burns with a yellow flame. After the burning the spoon is covered with a white layer (sodium rust).

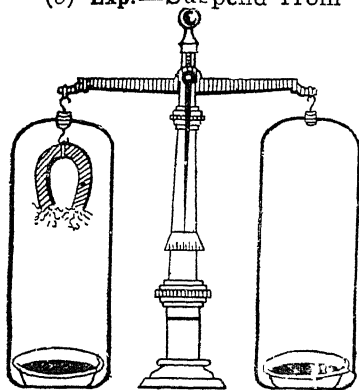
Magnesium burns with a dazzling white flame, leaving behind a white powder (magnesium rust). Some of the powder escapes as a white smoke.

(d) A rectangular piece of bright copper foil is closely doubled and each of the three open sides is closed by bending the edges and pressing them flat. Then the whole is heated through and through so that it becomes red hot. If after cooling it is unfolded, the interior surface of the copper, unexposed to air, remains bright and unchanged.

A piece of bright magnesium ribbon, placed at the bottom of a crucible, is covered with a layer of clean dry sand and heated. Removing the sand after cooling, the magnesium is found unchanged.

**Even on heating metals do not rust, if air is excluded from them.**

(e) **Exp.**—Suspend from the hock of the left arm



of a balance a horse-shoe magnet, whose poles carry tufts of iron powder (filings, wool) and counterpoise the whole. Bring the flame of a spirit lamp just in contact with the suspended iron powder. The powder begins to glow. Remove the flame; the glow continues and spreads. At the same time the sinking of the beam on the side of the magnet shows an increase of

Fig. 13.

mass. The powder weighs now more than in the beginning.

**Exp.**—In a hard glass tube put some magnesium powder. Close one end loosely with some asbestos wool and the other end with a cork fitted with a glass tube. Weigh the whole tube. Heat for a time

the glass tube at the place where the magnesium powder lies. Pass through the glass tube a current of air. The magnesium begins to burn. The asbestos holds back the white smoke. After all the magnesium is burnt, allow the tube to cool. Weigh the whole again. There is an increase of mass (weight).

**The rust (or ash) of a metal weighs more than the metal itself.**

All these experiments lead to the following results :

1. Metals undergo a change in substance in the air.
2. The rusting of metals in air is a chemical change.
3. This chemical change takes place slowly at ordinary temperatures, but swifter at higher temperatures.
4. The spontaneous burning of sodium with a yellow flame and the burning of magnesium on kindling prove that in rusting heat, often light, are produced.
5. The products of the rusting of metals (the rusts) are chemical compounds, as they have greater masses than the masses of the metals consumed.

We therefore conclude that some invisible stuff from the air combines with the metals to form their rusts.

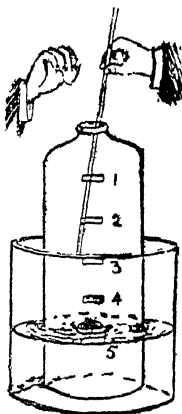
---

## VIII.

### THE COMPOSITION OF THE AIR.

1. **Composition of the air.**—The increase in the mass (weight) of metals, when exposed to the air or heated in it, shows that something from the air has combined with them. The question now is: Does the air as a whole, or only a part of it, combine with the metals? To answer this question it is necessary to make these substances undergo change in a limited volume of air. The volume must naturally become smaller, when combination occurs. Were the air an element, any residual gas left behind would exhibit the ordinary properties of air, such as the burning of a candle.

**Exp.**—On a light tin dish put some grams of iron powder (*ferrum limatum*). Place the dish with the iron powder on a piece of wood or cork floating on water in a glass trough. Take a large glass bell-jar (a large glass bottle with the bottom cut off will do) and close its neck with a tightly fitting stopper. Invert it and pour into it five tumblers of water, marking by means of a black line the level of the water in the jar, as each tumbler of water is poured into it. Then empty the jar. Cover the float in the trough with the bell-jar with its neck open (Fig. 14). Adjust the level of water in the trough till the water reaches the lowest mark



on the jar. Ignite the iron powder by touching it with the glowing end of an iron wire and close the jar with the stopper. Wait for some time till the glowing of the iron powder ceases and it cools. Observe that the water within the bell-jar rises, showing

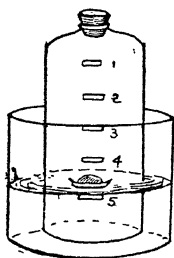


Fig. 15.  
jar open for a few minutes, and then plunge a burning

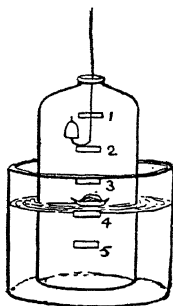


Fig. 16.

that a part of the air in the jar has been consumed in the process of burning (Fig. 15). Pour water into the trough till the level of water outside and inside the jar is the same. Observe that the water stands at the fourth mark, showing that one-fifth of the volume of the air has been consumed. Then remove the stopper and quickly introduce a burning taper fastened to a wire (Fig. 16). Observe that it is at once extinguished, showing that the residual gas is not ordinary air. Leave the

The same experiment may be performed more quickly as follows:—Instead of a jar employ a wide glass tube, about 1 metre long, closed at one end and fitted at the other end with a good stopper to which is attached a spoon of sheet tin. The stopper is taken out and the spoon filled with iron powder and heated in a flame till the powder glows. Keeping the tube horizontal

insert the stopper into the tube and press it tight. By shaking the tube vigorously, the iron powder becomes strewn in the whole tube and is brought into close contact with the air in the whole tube. On opening the tube under water the water rises into the tube, and on making the level of water inside and outside the same the water fills one-fifth of the tube (Fig. 17.)

Instead of iron a dry piece of phosphorus may be introduced into the tube and the tube stoppered. On



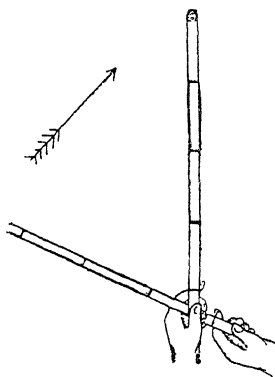


Fig. 17

closed with the thumb, to prevent the water from falling it is inverted. On removing the thumb and introducing a lighted match into the tube, the match goes out.

gently warming the phosphorus, it melts and takes fire. The burning phosphorus is made to run along the tube by a jerk the tube gets filled with white fumes. On cooling the stoppered end is placed under water and the stopper removed. The water rises, and on making the level the same inside and outside the tube, one fifth of the volume of air is found to have been consumed by the phosphorus in burning. The tube is then lifted out of the water and the open end being

**Exp.**—The following experiment with copper shows more clearly that only one-fifth of the volume of the air is active in the rusting of metals.

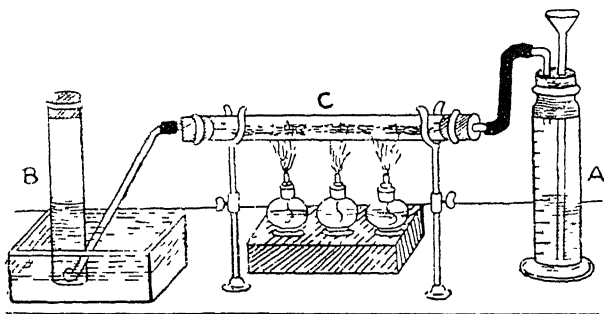


Fig. 18

Fill a hard glass tube C (a combustion tube) with

bright copper gauze rolled up, and connect it with equal measuring bottles A and B (Fig. 18). The bottle A is filled with air before the experiment, while B contains water. Pour water through the funnel. The air in A is pushed through the tube C into B. When A is wholly filled with water, B gets filled with air. Empty A and fill B with water, and arrange the apparatus as before. Heat the tube C gently at first and then more strongly. Notice that the copper undergoes almost no change. Now gradually fill A with water so that bubbles of gas rise slowly in B. Observe that the copper begins to glow, and that when all the air in A is displaced, only four fifths of B is filled with a gas, the remaining one-fifth being still occupied by water. Remove the delivery tube from the trough, and cool the combustion tube. Observe that the bright red copper has become black.

It may be asked whether the copper may not consume a greater quantity of air. But experiment proves that neither by stronger heating, nor by employing a larger quantity of copper, nor by passing a larger quantity of air it is possible to obtain any other result. More than one-fifth of the volume of air is never consumed, and four-fifths of the volume always remains behind.

By passing a large volume of air over heated copper collect three jars of the residual gas.

(a) Introduce a lighted taper into one of the jars. Notice that the candle goes out at once.

(b) Add lime water to another jar. Observe that the lime water remains clear, showing that the gas is not carbonic acid gas.

(c) Put a small mouse into the third. Observe that it is choked to death.

This colourless, inodorous gas, which remains behind after a metal rusts in a confined volume of air, is

lighter than air. One litre of it at  $0^{\circ}$  C. and 76 cm. pressure weighs 1.257 grams. It neither burns nor allows a candle to burn in it. Nor does it support life. It does not turn lime water milky. This inactive part of the air consists almost wholly of an element called **nitrogen**. Besides nitrogen it contains an extremely small percentage of other gases (argon, xenon, krypton, &c).

The name *nitrogen* means nitre generator, and is due to its being a component of nitre or saltpetre.

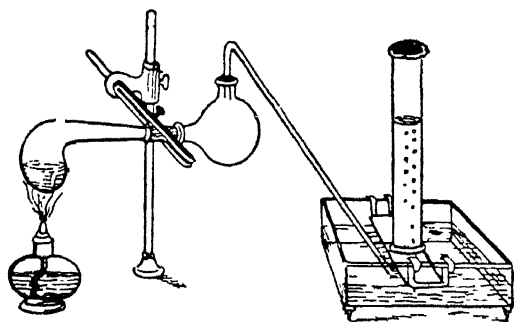
---

## IX

### THE COMPOSITION OF THE AIR.

If the inactive part of the air consists almost wholly of nitrogen, what is its active part? The simplest means of finding an answer to this question would be to decompose directly the rust or ash obtained by heating a metal in air. Now one of the methods of decomposing a compound is to heat it strongly. But the rust or ash of most metals is not decomposed by heat. Exceptions to this are red lead and the red ash obtained by heating mercury in an open dish for a long time at a temperature close to its boiling point.

**Exp**—Fill a small retort of hard glass nearly to a third with red mercury ash, and fit its neck into a tubulated receiver from which proceeds a delivery tube to a *pneumatic* or gas-collecting trough. Heat



the red ash at first gently, then strongly. After some time the ash turns dark red and finally dark brown. Observe the formation of a mirror of mercury in

Fig. 19.

the neck of the retort and the gradual collection of mercury in the receiver. At the same time bubbles of gas rise in the jar standing over the opening of the delivery tube in the trough. When the jar is full, remove the delivery tube from the trough and then remove the flame. Introduce a glowing splinter into

the jar. Observe that the splinter bursts into flame. As we have already seen, the rekindling of a glowing splinter is a test for the element oxygen. Hence the gas evolved by the red mercury ash when strongly heated is oxygen. As the red ash of mercury was originally obtained by heating metallic mercury, an element, in the air, the element oxygen must be present in the air. **Oxygen** is therefore the active part of the air.

Also red-lead when strongly heated gives off oxygen, but leaves behind not the metal lead but litharge.

Hence the air consists almost wholly of two elements, nitrogen and oxygen, and that about  $\frac{4}{5}$  of the volume of air is nitrogen and the remaining  $\frac{1}{5}$  is oxygen.

We have still to find out whether the air is a chemical compound or a mechanical mixture of oxygen and nitrogen.

**Exp**—Attach a candle to the cup of a deflagrating spoon, and after lighting it introduce it into a glass jar and close the jar with the metal plate attached to the spoon. With the aid of a watch count the number of seconds during which it burns in the jar. Find the volume of the jar with a measuring cylinder, and divide the jar into five equal parts by marks on the outside. Fill over water four-fifths of the jar with nitrogen obtained by passing air over heated copper and the remaining one-fifth with oxygen obtained by heating the red ash of mercury. Close the mouth of the jar with a glass plate and remove it from the trough. Invert the jar, remove the plate, introduce as before a burning candle on a deflagrating spoon and quickly close the mouth of the jar with the plate of the spoon. Observe that the candle burns for the same length of time in the artificial mixture of nitrogen and oxygen as it does in ordinary air.

(a) A litre of this artificial mixture has just the same mass as a litre of ordinary air. The mass of 1 litre of nitrogen is 1.257 grams and that of one litre of oxygen is 1.429 grams. Therefore the mass of one litre of the

mixture is  $(\frac{4}{5} \times 1.257 + \frac{1}{5} \times 1.429)$  or 1.291 grams, which is very nearly equal to that found by experiment for ordinary air.

(b) Oxygen vigorously supports the burning of a candle, while nitrogen does not at all support it. The air exhibits the properties of both to some extent, and its properties are the mean of those of its constituents. But the properties of a chemical compound are never the mean of its components.

(c) When liquefied air is allowed to evaporate, nitrogen goes off first, leaving liquid oxygen behind. Such a separation of components never takes place in a chemical compound.

(d) Were the air a chemical compound, the air dissolved in water would also contain oxygen and nitrogen in the same proportion by volume as ordinary air. But the air dissolved in water is much richer in oxygen than ordinary air. In 100 parts by volume of air dissolved in water there are 34 parts of oxygen and 66 of nitrogen.

(e) When oxygen and nitrogen are brought together to form air, there are no indications of a chemical combination, such as change of temperature, volume, &c.

All these facts establish that the atmospheric air is a mixture of oxygen and nitrogen and not a chemical compound, although oxygen and nitrogen are found in nearly the same proportion by volume (21 % of oxygen and 79 % of nitrogen) in all parts of the globe.

Oxygen was first prepared by Scheele, a Swedish chemist, and about the same time, independently of Scheele, Priestley obtained it by heating the red ash of mercury. But Lavoisier was the first to give a correct explanation of the rusting or calcination of metals when heated in air. He placed mercury in a glass retort the neck of which, drawn out and bent, passed under mercury in a glass trough and then to the air contained in a bell-jar (Fig. 20). The mercury in the

retort was heated by means of a furnace for many days. As this went on, red specks appeared on the surface of the mercury in the retort, while the mercury gradually rose in the jar indicating a decrease of volume of the air in it. The heating was stopped after twelve days and the decrease of volume of the air in the jar was noted. The red ash was then collected and introduced into a glass tube closed at one end. The end

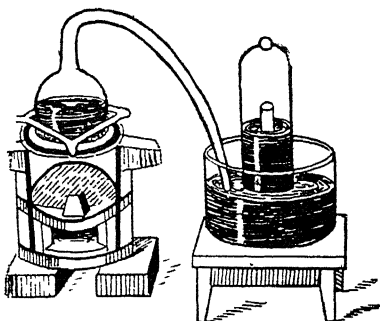


Fig. 20

of this tube was heated, and the gas evolved was collected in a graduated cylinder over mercury. The volume of the gas was found to be exactly the same as the decrease of volume of air in the bell-jar. Lavoisier found that this gas supported vigorously the burning of a taper, while the residual gas in the bell-jar extinguished a burning candle and quickly suffocated a mouse. He called the former oxygen and the latter **azote** (without life).

Lavoisier's experiments clearly established for the first time that there was no loss of substance in burning, and that the gain in mass by metals in rusting was exactly equal to the loss sustained by the air.

The products of the transformation of metals in air are in general compounds of metals with oxygen and are called **oxides**. The oxide of iron is reddish-brown, the oxide of copper is black, the oxide of tin is yellowish white, the oxide of lead is yellow or red, the oxide of magnesium is white, the oxide of mercury is brick red, the oxide of zinc is white. The combination of metals with oxygen is called **oxidation**.

The oxides of many metals occur in nature as **ores**. Gold, platinum, silver are called the **noble metals**, because they are not chemically affected by the oxygen of the air. Mercury and aluminium form no oxides at ordinary temperature and are named the semi-noble metals. All other metals are called the **base metals**. Why?

*Question.*—Lavcisiert partly filled a glass retort with metallic tin and completely sealed it. The retort was then weighed. By heating the retort the metal lost its brightness and underwent a change. On cooling the retort was weighed again and it was found to weigh the same as before. But on opening the sealed end of the retort and weighing it again it was found to weigh more. What inference do you draw? What would have happened, if the sealed end of the retort had been opened under water? And why?

---



# Syllabus.

## X

### OXYGEN.

1. **Preparation of oxygen.**—To study the properties of oxygen large quantities of the gas will be needed. To prepare a large quantity of oxygen from the red oxide of mercury will be expensive. Consequently we resort to other substances, which are not only cheaper but also give out more oxygen more easily. One such substance is **potassium chlorate**, an artificial product, which is sold in small white crystalline plates under the name of **chlorate of potash**. You made use of this compound in an experiment while studying chemical decomposition. 5 grams of potassium chlorate yield about 1.4 litres of oxygen, whereas 5 grams of the red oxide of mercury yield not more than 0.25 of a litre.

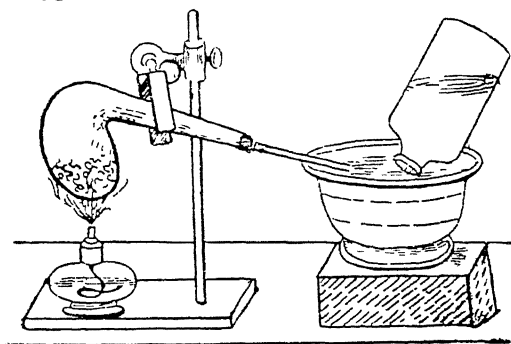
**Exp.**—Heat a small quantity of potassium chlorate in a test tube. Observe that it first melts, and then gives off bubbles of gas. Introduce a glowing splinter; it bursts into flame indicating that the gas is oxygen. The residue in the tube is a white solid, whose properties, as you have already seen, are different from those of the chlorate. Potassium chlorate is a compound of the elements, potassium, oxygen and chlorine, whereas the residue consists only of potassium and chlorine.

**Exp.**—Heat some potassium chlorate in a flask a little above its melting point just till no gas is given off as tested by a glowing splinter. Remove the flame and immediately drop a pinch of powdered manganese dioxide. Observe the vigorous evolution of oxygen. Allow the flask to cool, boil the stuff in the flask with water and filter the solution. The filtrate contains potassium chloride, as can be found by the tests you know, while the residue on the filter is unchanged manganese dioxide, which can be used again.

A substance, such as manganese dioxide, which by its presence hastens a chemical reaction without itself suffering a chemical change, is called a **catalyst** and its action is called **catalytic action**.

Impure manganese dioxide often contains inflammable substances (charcoal, &c.); then on heating the mixture sparks appear. If these impurities are present in large quantity, a dangerous explosion may be produced. It is therefore advisable to heat the dioxide to redness before using it. The weight of manganese dioxide to be added is about one-fourth of the weight of the chlorate.

**Exp.**—Introduce a mixture of 120 grams of potassium chlorate and 30 grams of manganese dioxide (**oxygen mixture**) into a glass retort and fit up the apparatus, as shown in Fig. 21. To



avoid breakage an iron retort or a copper flask may be employed. Heat gently, and allow the first bubbles to escape, as these consist of air from the vessel.

Fig. 21,

Collect seven or eight jars over water. When the collecting is finished, first remove the delivery tube from the water and then take away the flame. Why? Clean the retort soon after the experiment with hot water.

Oxygen is sold compressed in cylinders of steel, just like hydrogen. It is prepared by the electrolysis of water or by the liquefaction of air. As already pointed out, liquid nitrogen evaporates much more rapidly than liquid oxygen. So when liquid air is evaporated in

apparatus specially constructed for the purpose, nitrogen evaporates, leaving behind liquid oxygen.

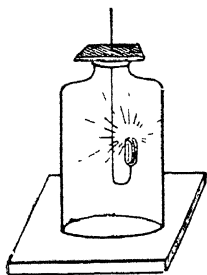
Barium oxide, when heated to  $700^{\circ}\text{C}$  in air under pressure, absorbs oxygen and forms a higher oxide. On allowing the nitrogen to escape and reducing the pressure the higher oxide changes again into the lower oxide, giving off oxygen. This method has been employed for obtaining oxygen from the air.

**2. Properties of oxygen.**—Oxygen is a colourless, inodorous, and tasteless gas. It is slightly denser than air; 1 litre of oxygen has a mass of 1.43 grams, while the mass of 1 litre of air is 1.293 grams. It liquefies below  $-118^{\circ}\text{C}$ ; liquid oxygen is blue.

The most important chemical property of oxygen is that many substances **burn** in it, that is to say, it combines with them evolving sufficient heat to render them incandescent. We say that **oxygen is a supporter of combustion**.

To examine the products of combustion perform the following experiments:—

**Exp.—(a)** Attach a piece of charcoal (3 cm. long and 1 cm. thick) to the bent end of a wire (Fig. 22), hold it in a flame till it glows, and then introduce it into a jar of oxygen. The charcoal glows very intensely and burns with a dazzling light. After the burning has ceased, pour some distilled water into the jar. Cover the jar with a glass plate and shake it well. To a portion of this solution, add some solution of **blue litmus, which is a reagent for acid substances**. The reddening of blue litmus is



**Fig. 22.** an acid reaction. Observe that the solution turns red, showing that the aqueous solution of the colourless, inodorous, gaseous product of the burning of charcoal in oxygen is acid in reaction. To the remainder of the solution add clear lime water. Observe that the lime

water turns milky, indicating the presence of carbonic acid gas in solution.

Wood charcoal contains the element carbon, but it is not pure carbon. When burnt, it leaves ash behind. Try. While burning in oxygen the carbon of the charcoal combines with the oxygen, and produces carbonic acid gas or **carbon dioxide**, which turns lime water turbid. Carbon dioxide is a compound of the elements, carbon and oxygen.

Carbon + oxygen = carbon dioxide.

(b) Ignite a piece of sulphur in a deflagrating spoon. Observe the faint blue flame with which sulphur burns in air. Now plunge the burning sulphur into a jar of oxygen (Fig. 23). Observe its more vigorous burning in oxygen producing white fumes of a suffocating odour (smell of burning sulphur).

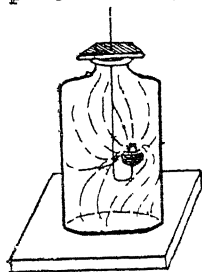


Fig. 23.

Remove the spoon, and distilled water, close the mouth of the jar with the palm and shake vigorously. You feel a pressure on the hand owing to the partial vacuum produced by the solution of the gas in water. The gas is very soluble in water, and is therefore absorbed by the water when the jar is shaken.

Add to a portion of this solution a little blue litmus solution. The litmus turns red, indicating that the solution of the gas is acid.

To the remainder of the solution add lime water. No change is perceived. What does it show?

When burning the element sulphur combines with the element oxygen to form the gaseous compound **sulphur dioxide**, which has a suffocating odour and is highly soluble in water. Sulphur dioxide is a compound of the elements, sulphur and oxygen.

Sulphur + oxygen = sulphur dioxide.

(c) Cut a bit of phosphorus of the size of a pea,

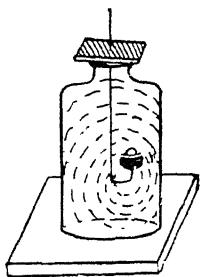


Fig. 24.

quickly dry it by passing it between folds of filter paper, place it in a deflagrating spoon, and warm it. As soon as it begins to burn, plunge it into a jar of oxygen (Fig. 24). Observe the phosphorus burning in oxygen with a dazzling white light evolving dense white fumes. When the burning has ceased, remove the spoon, pour a little water into the jar and shake well. The fumes rapidly dissolve in water. How do you know? Note that this solution also turns blue litmus red, indicating that this solution also is acid in reaction. The white fumes consist of fine particles of **phosphoric oxide**, a compound of the elements, oxygen and phosphorus.

Phosphorus + oxygen = phosphoric oxide.

(d) Tie one end of a spiral of magnesium ribbon or wire to a deflagrating spoon, ignite the other end, and introduce it into a jar of oxygen. It burns with an extremely dazzling light, and is transformed into a white powder, which is feebly soluble in water. The moistened powder turns red litmus blue, showing the alkaline character of the powder produced. The white powder is a compound of the element magnesium with oxygen. Magnesium + oxygen = magnesium oxide.

(e) Place a small dry piece of sodium on a deflagrating spoon, heat it over a flame till it begins to burn, and then plunge the spoon into a jar of oxygen. The sodium burns with a bright yellow flame, leaving behind on the spoon a white residue. After cooling, add a little water to the spoon; the residue rapidly dissolves. Wet your fingers with the solution, and note the slimy or soapy feeling and the caustic action on the skin. Add more water. To one portion of the solution add a little blue litmus solution, and to another portion add a little red litmus solution. No change is perceived in the former, whereas the latter turns blue, showing the alkaline nature of the compound of the element sodium with oxygen.

Sodium + oxygen = oxide of sodium.

The oxide of sodium combines with water to form **caustic soda**.

(f) Make a spiral of fine iron wire, fix the tip of a sulphur match to one of its ends, and attach the other end to a piece of card-board. Light the match and introduce the wire into a jar of oxygen (Fig 25), which contains some water. The match burns and ignites the iron. The iron continues to burn, throwing off brilliant sparks, while drops of a molten stuff fall into the water. Observe that this molten stuff does not dissolve in water and has no action on blue or red litmus. A substance

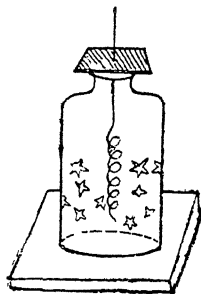


Fig. 25. which has no action on either blue or red litmus is said to have a **neutral reaction**. The compound of iron with oxygen is neutral in reaction.

Iron + oxygen = oxide of iron

The products yielded by the burning of carbon, sulphur, phosphorus, sodium, magnesium, iron in oxygen are chemical compounds, called oxides. The aqueous solutions of some of these oxides, called also **anhydrides**, have an acid(sour) taste and turn blue litmus red. These solutions are called **acids**; whence the name oxygen, which means acid generator; but all acids do not contain oxygen. The solutions of other oxides have the taste of lye and turn red litmus blue; they are called **alkalies**.

A very active form of oxygen, which oxidises most metals at ordinary temperatures and attacks organic substances, is the strong-smelling **ozone**. As it destroys disease-spreading organisms, it is employed to purify air and sterilise water. It is produced by subjecting oxygen to the electric discharge.

Oxygen is employed to produce very hot flames, such as the oxy-hydrogen and oxy-acetylene flames. No animal can live without the oxygen present in the air. In cases of collapse and difficult breathing pure oxygen is substituted for the air.

## XI

### WATER.

1. **The three states of water.**—Water occurs in nature in all the three physical states, in the gaseous state as **steam**, in the liquid state as **water**, in the solid state as **ice**.

**Exp.**—Bring into a room a glass tumbler filled with ice-cold water and well dried on the outside. The surface of the glass becomes dim. The dimness is due to the deposit of fine droplets of water on the cold sides of the glass by the condensation, that is, liquefaction, of the aqueous vapour present in the air coming in contact with the glass.

Evaporation is constantly taking place from the surfaces of seas, lakes and rivers, from the moist earth, from plants, from the bodies of animals. Water vapour is also given to the air by certain substances burning in air. A burning candle, for example, gives off water vapour.

The mass of aqueous vapour in a cubic metre of the air may vary largely in different places and at different times. But some water vapour is always present. Cold dry air contains a less percentage of aqueous vapour than warm damp air. The water vapour in the air is continually condensing and appears as clouds, mist, fog, dew, rain, snow, hail and frost.

In the liquid state water occurs in nature in enormous quantity. About three-fourths of the surface of the earth is covered with water. The soil and porous rocks contain considerable quantities of water. Plants and animals contain a large proportion of water.

**Exp.**—Heat in a dry test-tube one after another a small piece of wood, meat, potato, or any fresh

vegetable. Hold the open end of the test-tube slightly lower than the closed end. Observe the drops of water falling from the test-tube.

Most animal and vegetable substances behave similarly. The human body is nearly 70 per cent water. No life can subsist without water.

Water in the form of solid ice covers the coldest parts of the surface of the globe, namely, the polar regions and the tops of high mountains.

**2. Pure water.**—One of the most important properties of water is its great solvent action on a large number of substances. Owing to this property it is very difficult to obtain water in a pure state. Spring water, for example, contains various gases or solids in solution.

**Exp.**—Heat water in a glass flask so that you

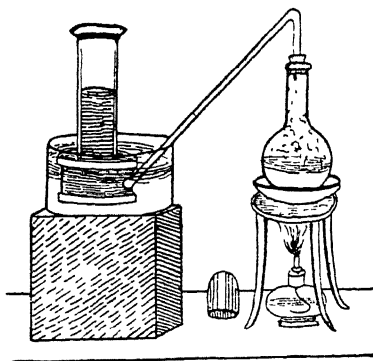


Fig. 26.

can observe what is taking place in its interior. We soon see in the middle of the liquid fine bubbles of gas which rise to the surface and disappear. It is the dissolved gas that escapes. The gas may be easily collected (Fig. 26). In the figure the glass flask is filled to the brim with tap water and the delivery tube also is filled with water so as to exclude all air. On boiling

the water the escaping gas will collect in the jar. Steam will be condensed by the water in the trough.

**Exp.**—Evaporate to dryness some cubic centimetres of tap water in an evaporating dish supported on a wire gauze. Observe a grey deposit of solid matter left behind.



No water found in nature is pure. Even rain water contains gases and dust washed from the air.

Rain water, falling on the ground, takes up impurities. Some of the water flows along the surface, becomes more and more impure, and finally reaches the ocean. The remainder soaks into the ground, percolates through the soil, and finally finds its way again to the surface as a spring or well, or collects in a lake, or flows from a hill-side along a river. Consequently the waters of springs, lakes and rivers contain considerable quantities of dissolved substances. If the dissolved matter in spring water gives it medicinal properties, the water is called a **mineral water**.

**Sea water** contains about 3·6 per cent by weight of dissolved solids of which 2·6% is common salt. Some inland seas, like the Dead Sea, contain a much larger proportion of dissolved solids.

A water which, like rain or distilled water, lathers readily with soap, is said to be **soft**.

**Exp.**—Prepare some soap solution by dissolving slices of pure soap (such as Brown Windsor Soap) in hot distilled water. Pour a few drops of this solution in distilled water and shake. Observe that a lather is easily formed.

A water which does not immediately lather with soap is said to be **hard**. When soap is added to a hard water, a white curdy precipitate is formed, which floats on the surface. The hardness of water is due to the presence of dissolved calcium and magnesium compounds. Soap produces with magnesium and calcium compounds present in the water insoluble compounds, which are therefore precipitated.

**Exp.**—Add some lime water to tap water, and then add some soap solution. Note that a curd or scum is formed on the water instead of a lather. Hard water is unfit for cooking legumes (dhal), as the calcium compounds present in them form insoluble compounds with the substances present in the seeds. Hard water is unsuited for laundering, as it consumes

much soap. Hard water covers the inside of kettles and boilers with an incrustation or fur, which increases the amount of fuel required to heat the water. The turbidity produced by boiling and the absence of a ready lather with soap solution serve to recognise hard waters. Certain substances, such as ammonia, lime, washing soda, or borax, soften hard waters. Try.

Pure water is obtained by distillation. Distillation

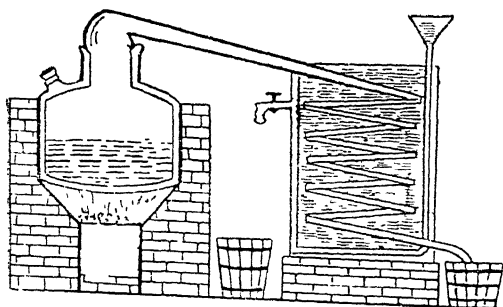


Fig. 27.

is effected in a condenser or still (Fig. 27). Water is converted into steam and is condensed in tin pipes. The first portion of the distillate is rejected, as it contains the gases or volatile liquids

dissolved in the impure water.

**Exp.**—Test the purity of distilled water by slowly evaporating a few drops on a watch glass in a room free from dust. There should be no stain left on the glass.

Suspended impurities in water are removed by filtration. Sand filters are used in many cities for purifying muddy water. Muddy water is sometimes treated with alum, which makes the minute particles of mud to come together and form larger particles which sink to the bottom.

**3. Properties of pure water.**—Pure water is an inodorous, tasteless, colourless liquid. When we look through a long column of the liquid, it has a blue or greenish blue colour. One cubic centimetre of water at  $4^{\circ}\text{C}$ . has a mass of 1 gram. Above and below  $4^{\circ}\text{C}$ . one gram of water occupies a larger volume than one cubic centi-



## XII

### THE COMPOSITION OF WATER.

1. **Action of metals on water.**—Many of those metals which undergo oxidation in air, either at ordinary temperature or when heated, act upon water and liberate a gas.

Such metals as sodium, potassium, calcium, act upon water at ordinary temperatures.

The metals, magnesium, zinc, aluminium, iron, react with water only when heated.

Copper, mercury, lead have no action upon pure water at any temperature.

**Exp.**—Sodium is a silver white soft metal of specific gravity 0.97. It oxidises very rapidly in the air, forming a white solid which rapidly absorbs moisture. It is therefore preserved under petroleum. When heated in air, it burns with a yellow flame.

Place a bit of sodium on water. Observe that it changes into a round ball, and moves to and fro on the water with a hissing sound, until it finally disappears. The hissing implies the evolution of a gas.

To collect this gas, wrap a piece of sodium of the size of a small pea in a strip of filter paper and introduce it by means of forceps under an inverted jar standing in a trough containing water (Fig. 28). The sodium separates itself from the paper and rises in the jar, while at the same time a gas collects in it. When the jar is full of gas, take it out of the water still holding it in an inverted position. Note that the gas is colourless, odourless, lighter than air. Apply a lighted taper to its mouth. The gas burns with a pale blue flame. Thrust

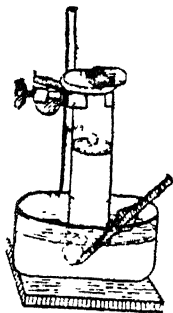


Fig. 28.

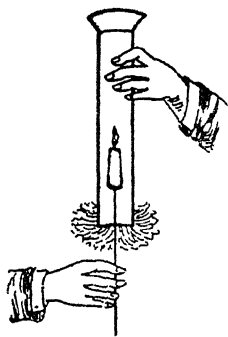


Fig. 29.

the taper inside; it is extinguished (Fig. 29). Draw out the taper to the burning gas, it is rekindled. Thrust it up into the jar again; it again goes out. This may be repeated a number of times. It therefore follows that the gas liberated from water by the chemical action of sodium on it burns in air but does not support the burning of a taper in it. The gas is neither oxygen nor nitrogen. This inflammable gas is the element **hydrogen**.

Sodium + water = Hydrogen + sodium oxide.

The sodium oxide formed dissolves at once in the water forming caustic soda and the water becomes alkaline. Prove this.

**Exp.**—Like sodium potassium is a silver white soft metal of specific gravity 0.86. It very rapidly oxidises in air into a white deliquescent solid. Potassium, like sodium, has to be preserved under petroleum.

Throw a bit of potassium on water. Note that it at once takes fire and burns with a violet flame; while at the same time it moves to and fro over the water. The heat generated by the chemical action between potassium and water is so great that the metal and the gas set free take fire. This metal is not easy to manipulate.

Potassium + water = Hydrogen + potassium oxide.

The potassium oxide formed dissolves in water forming caustic potash, which renders the water alkaline.

**Exp.**—Metallic calcium is a white metal, which can be cut with a knife, though not so soft as potassium or sodium. It slowly decomposes water at ordinary temperatures. Put some shavings of calcium in water in an evaporating dish and place over it an inverted test-tube filled with water. Test the gas.

Hydrogen is slowly evolved owing to the adherence on the metal of the slaked lime formed by the action of water on the metal.

Calcium + water  $\rightleftharpoons$  Hydrogen + slaked lime.

Slaked lime is a compound of calcium, hydrogen and oxygen.

**Exp.**—In a flask put some water and magnesium powder and heat the water to boiling. Allow the first bubbles of gas to escape and collect the remainder in a jar over water. Observe that the gas which collects is hydrogen. The oxide formed adheres to the metal and hinders further action.

**Exp.**—The chemical action of water on magnesium may be observed in another way. Fit up the apparatus shown in Fig. 30. Clean a strip of magnesium ribbon, roll it into a spiral and push it into a short combustion tube drawn out to a jet at one end. Fill one third of the flask with water and set it on wire gauze resting on the ring of a retort stand, and place the cork and tubes in position. Heat the water in the flask with a small flame, and at the same time keep the whole of the tube containing the magnesium hot by moving a flame-

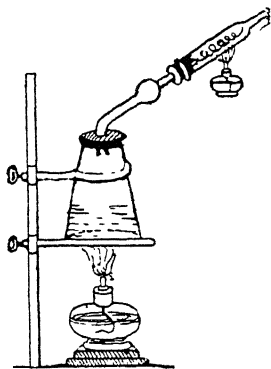


Fig. 30

backwards and forwards over its whole length. This prevents the steam from condensing in the combustion tube, while the bulb prevents drops of water from entering the tube. As soon as a rapid current of steam escapes, heat the magnesium strongly till it begins to glow; then stop heating the tube and apply the flame to the open end of the tube. The gas escaping at the jet burns in air. Heated magnesium burns with a dazzling light in the same manner in vapour of water as it does in air. The product of the burning in steam is a

white solid resembling the oxide of magnesium produced by its burning in air. Hence water must contain both hydrogen and oxygen.

Magnesium + water = Hydrogen + oxide of magnesium.

**Exp.**—Allow iron to lie in contact with ordinary water exposed to air. After some days the iron is found covered with rust.

Burnish some iron tacks with emery paper and place them in water which has been boiled for some time to expel all dissolved air, and tightly cork the vessel. No rusting occurs even after many days. Boil some iron filings with water. Note that there is no apparent change. These experiments show that both water and air are needed for the rusting of iron at ordinary temperature. The brown rust is not a simple oxide containing only iron and oxygen.

**Exp.**—Conduct steam through a red-hot iron tube

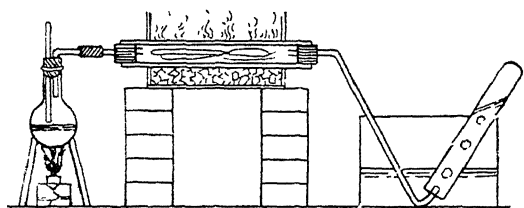


Fig. 31

filled with bright iron wires (or tacks) and collect a number of jars of the gas which escapes by placing the

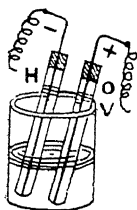
end of the delivery tube in a trough of water (Fig. 31). When the tube has cooled, pour out its contents on a piece of paper and compare them with the iron wires. Observe that the iron wires have lost their brightness and become black. Had the tube with the iron wires been weighed before the experiment, it would be found to have increased in mass after the experiment. The black oxide looks precisely like the one previously obtained by burning iron in air or oxygen.

Iron + water = Hydrogen + oxide of iron.

The action of metals on water shows that water is a compound of an inflammable gas, hydrogen, and another which we suspect to be oxygen.

2. **Decomposition of water by the electric current (or electrolysis).**—By the action of various metals on water, the existence of hydrogen in it is proved. The formation of the oxides of magnesium and iron, when steam is passed over glowing magnesium and iron, indicates presence of oxygen in water. But this proof is indirect. Although it is not easy to separate the elements of water from one another by the application of heat, we can separate them by an electric current.

**Exp.**—Pure water does not conduct the electric current. Put distilled water in the apparatus shown in Fig. 32. V is a glass trough containing distilled



water; H and O are glass tubes (2 cm. diameter) fitted with two rubber stoppers through which pass bright thin iron rods (called *electrodes*.) These rods carry electricity through the liquid. The tubes H and O are completely filled with distilled water and inverted into the trough. The apparatus is called a **voltmeter**. Connect the iron rods with the two **poles** of a

Fig. 32. **voltaic battery** (a set of two Daniels cells used by electroplaters) or an **accumulator**. Nothing in particular happens. But add to the water in the trough some drops of a solution of caustic soda, which enables the water to conduct the current. At once bubbles of gas make their appearance along the *electrodes* and collect in the tubes. Observe that the tube H connected with the negative pole of the battery is filled much more rapidly, stop the experiment. You observe:

1. That the volume of gas collected in the tube H (connected with the *negative pole* of the battery) is exactly double the volume of gas collected in the tube O (connected with the *positive pole*)

2. That the gas of the tube H burns on lifting the tube out and applying a flame, just in the same manner



as the gas produced by the action of sodium. calcium magnesium, or iron upon water

3. That the gas of the tube O kindles a red hot splinter and is therefore oxygen.

It can be shown that the solution of caustic soda added for rendering the water conducting suffers no change, so that the evolution of the two gases is produced by the decomposition of the water alone.

3. **There are in water only oxygen and hydrogen.**—The preceding experiment shows that from water we can get oxygen and hydrogen, further that they are the only components of water. We shall see later that it is possible to produce water which is exactly identical with the purest distilled water by starting with oxygen and hydrogen as original substances. So we can say that water is a compound of hydrogen and oxygen and name it an **oxide of hydrogen**.

---

## XIII

### HYDROGEN.

#### 1. Preparation of hydrogen. Exp.—Fit a flask (Fig. 33)

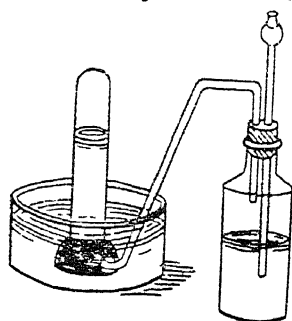


Fig. 33.

with a stopper with two holes. Pass a thistle funnel through one hole and a right-angled bent tube through the other. Introduce into the flask small strips of sheet zinc, and close it with the stopper, seeing that the end of the tube of the funnel reaches right down to the bottom of the flask. By means of a short piece of rubber tubing connect the right-angled tube with a delivery tube passing under the bee-hive shelf of a pneumatic trough.

Prepare dilute sulphuric acid (1 part by volume to 10 of water) by pouring the acid in small quantities into water and stirring with a glass rod. The vessel in which the mixture is made must be thin-walled and cooled by being placed in water, as there is considerable heating when concentrated sulphuric acid is poured into water. Hence great care is needed in mixing water and sulphuric acid. Pour the dilute sulphuric acid through the funnel until the zinc is covered. The zinc dissolves in the acid and hydrogen gas is given out.

In the place of zinc we can substitute iron or magnesium, and instead of dilute sulphuric acid dilute hydrochloric acid can be employed. All acids contain hydrogen and in general this hydrogen is liberated by the action of metals on dilute acids. Before Lavoisier hydrogen was called *inflammable air*.

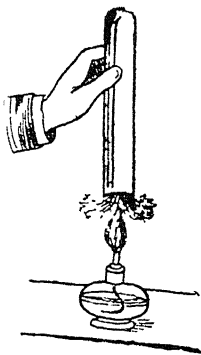


Fig. 34. Collect several jars of gas in this way.

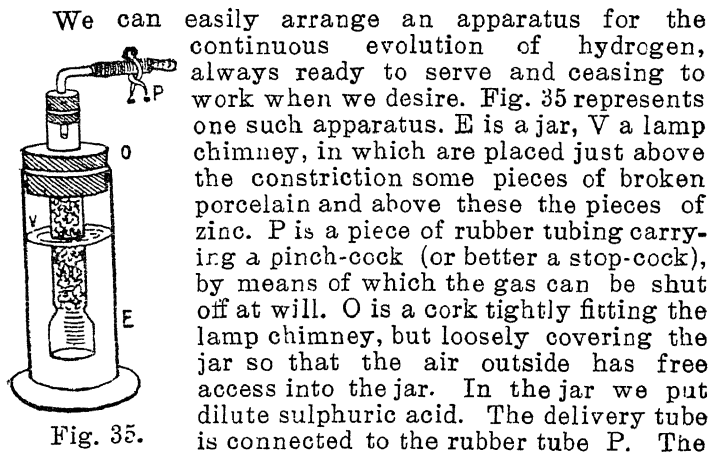


Fig. 35. We can easily arrange an apparatus for the continuous evolution of hydrogen, always ready to serve and ceasing to work when we desire. Fig. 35 represents one such apparatus. E is a jar, V a lamp chimney, in which are placed just above the constriction some pieces of broken porcelain and above these the pieces of zinc. P is a piece of rubber tubing carrying a pinch-cock (or better a stop-cock), by means of which the gas can be shut off at will. O is a cork tightly fitting the lamp chimney, but loosely covering the jar so that the air outside has free access into the jar. In the jar we put dilute sulphuric acid. The delivery tube is connected to the rubber tube P. The air is rapidly driven out of the chimney by the hydrogen evolved, when the zinc comes in contact with the acid. If the apparatus is closed by means of the pinch-cock, the gas, produced for a moment, pushes out the acid which acts upon the zinc.

The evolution of gas ceases till the pinch-cock is opened.

What is it that remains in the hydrogen-generating apparatus, when the zinc has dissolved and disappeared?

**Exp.**—Take out some of the liquid from the apparatus and add zinc till no more will dissolve. Filter, if necessary, and evaporate the solution partially in a basin. On cooling transparent, colourless, prismatic crystals are formed. These crystals contain water of crysallization, and on heating they lose the water of crystallisation and fall to a white powder. This white powder is called **zinc sulphate** and is a compound of zinc with what remains of the compound sulphuric acid after the removal of hydrogen from it.

Sulphuric acid + zinc = hydrogen + zinc sulphate.

In commerce hydrogen, prepared on a large scale by the electrolysis of a solution of caustic soda, is sold compressed in steel cylinders. It is enough to open a stop-cock to obtain a regular current of the gas.

**2. Properties of hydrogen.**—Hydrogen is a colourless, inodorous and tasteless gas. It is 14·5 times lighter than ordinary air. A litre of hydrogen weighs only 89 milligrams.

**Exp.**—To prove the lightness of hydrogen place a jar of hydrogen mouth downwards over a jar full of air. Keeping the mouths of the jars in contact, turn them together upside down. It is found that the gas in the upper jar which originally contained air is now inflammable; it now contains hydrogen. Soap bubbles can be blown with hydrogen; these rise in the air and rapidly reach the ceiling of the room. These bubbles can be inflamed. **Hydrogen is the lightest of all known gases. Hydrogen is therefore employed to fill balloons.**

On account of its lightness dry hydrogen may be collected easily by the upward displacement of air (Fig. 36). To dry it hydrogen is made to bubble through strong sulphuric acid.

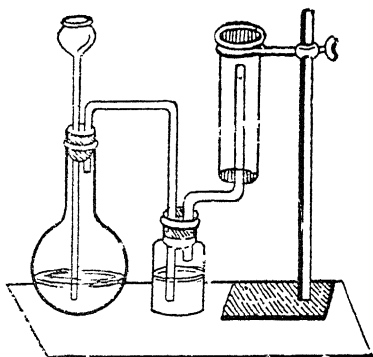


Fig. 36.

air solidifies at a higher temperature than the boiling point of hydrogen.

**Exp.**—You have already seen (page 59) that hydrogen burns but does not support combustion like oxygen. Hydrogen may be burnt at a jet (Fig. 37), but care must be taken to see that no air remains in the apparatus before the jet of gas is lighted.

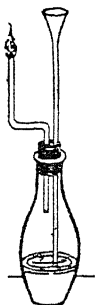
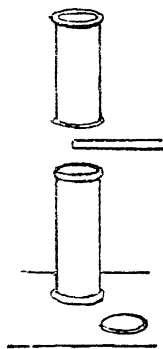


Fig. 37.

**Exp.**—Fill a dry jar with dry hydrogen by the upward displacement of air and place it mouth downwards over another dry jar containing air. Keeping the two jars tight together, reverse their positions. There is no chemical union but only a mechanical mixture between the gases. Raise the upper jar and bring a red hot iron wire below it (Fig. 38). There is a low explosion, and the sides of the jar become dim owing to the condensation of aqueous vapour. Repeat the experiment with oxygen instead of air; the result is the same. Chemical combination between the gases may be brought about also by an electric spark.

Place a jar of hydrogen with its mouth under water. Observe that there is no perceptible rise of water into the jar, showing that the gas does not appreciably dissolve in water.

Hydrogen liquefies at  $-234^{\circ}\text{C}$  and solidifies at  $-257^{\circ}\text{C}$ . The bottom of a test-tube is immediately filled with solid air, when it is plunged in liquid hydrogen. So



**Exp.**—The explosion is loudest when we mix two volumes of hydrogen with one volume of oxygen or five volumes of air. Fill a small bottle one-third with oxygen and two-thirds with hydrogen by the displacement of water. Wrap the bottle in wire-gauze or in folds of cloth and then apply a flame to the mouth. A mixture of two volumes of hydrogen and one volume of oxygen is known as **deto-nating gas**.

Hydrogen flame is very hot; the flame is hottest ( $2400^{\circ}\text{C}$ ), when hydrogen and

oxygen are present in the proportion of 2 to 1 by volume. This is done in the oxy-hydrogen flame. Platinum easily melts in this flame. A piece of quick lime held in this flame becomes white-hot without melting and emits a dazzling light, known popularly as lime-light.

**3. Formation of water in burning hydrogen**—When a mixture of hydrogen and air, or of hydrogen and oxygen, is exploded in a jar or test-tube, or when a dry tumbler is held over the flame of a burning jet of hydrogen, the inside of the vessel gets dim owing to the deposit of some dew. This dew may be collected and examined.

**Exp.**—Fit up the apparatus shown in Fig. 39. The hydrogen from the generating apparatus is conducted through a tube containing dry calcium chloride. Instead of a tumbler, bring the well-dried outside of a large porcelain basin containing cold water in contact with the flame. Drops of liquid soon form on the bottom of the basin and fall. The liquid collected is limpid, colourless, odourless and tasteless; is neither alkaline nor acid; leaves no residue on evaporation. One c.c. of it weighs one gram. It freezes at  $0^{\circ}\text{C}$ . and boils at  $100^{\circ}\text{C}$ . All these are the properties of pure water.



## XIV.

### REDUCTION. COMPOSITION OF WATER.

1. **Reduction.**—A compound which any element forms with oxygen is called an **oxide** of that element, and the process by which an oxide is formed is known as **oxidation**. Many elements undergo oxidation when heated in air or oxygen. While oxidation takes place easily, it is not so easy to remove oxygen from an oxide. **The withdrawal of oxygen from oxides or other compounds of oxygen is called reduction.** Reduction is the reverse of oxidation.

The removal of oxygen from an oxide is very rarely effected by heat alone. Red oxide of mercury (minium partially) is an exception. From other common oxides it is not so easy to remove oxygen. To withdraw the oxygen from them they have to be heated with some substance that has a stronger affinity for oxygen than the elements with which it is already combined. Such substances are called **reducing agents**. The metals sodium, potassium, calcium, magnesium and iron have acted as reducing agents in removing oxygen from water. By heating many metallic oxides with reducing agents (*e.g.*, carbon) it is possible to obtain pure metals. This is the process by which many metals are obtained from their ores by heating them with carbon.

Carbon and hydrogen are good examples of reducing agents.

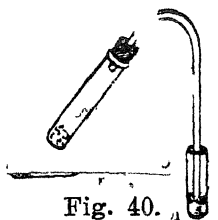
**Exp.**—Powder a little wood charcoal and mix it with the yellowish white oxide of tin. Heat the mixture strongly in a covered crucible. The charcoal takes the oxygen from the tin oxide and burns to carbonic acid gas, while metallic tin is left behind in the crucible.

Oxide of tin + carbon = tin + carbon dioxide.

**Exp.**—Mix a little powdered charcoal with about four times its weight of copper oxide, and introduce the



mixture into a test-tube of hard glass, fitted with a delivery tube (Fig. 40). Insert the end of the delivery tube into some lime water contained in a test-tube. Heat the mixture strongly. Observe that a gas bubbles up through the lime water, turning it milky, showing that the gas is carbon dioxide. Note the particles of red copper making their appearance in the heated test-tube. Any excess of charcoal from the residue in the ignition tube may be separated by washing.



Copper oxide + carbon = copper + carbon dioxide.

**Exp.**—The reducing character of hydrogen may be seen as follows. Hydrogen from a generator is dried by passing it through a tube containing calcium chloride and then conducted over copper oxide contained in a hard glass tube drawn out to a narrow end (Fig. 41).

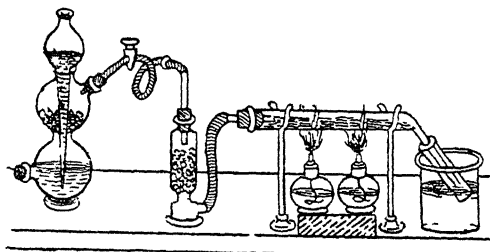


Fig. 41.

The narrow end of the hard glass tube is inserted in a dry test-tube which is cooled by water. The copper oxide undergoes no change at ordinary tempera-

ture. Gently heat the copper oxide. Observe the steam condensing in the test tube, while the black oxide of copper is turned into red metallic copper. Continue to heat until the whole of the oxide is changed. Remove the flame and allow hydrogen to pass through the tube until it is cold. In this case the hydrogen combines with the oxygen of the copper oxide and reduces the latter to metallic copper.

Copper oxide + hydrogen = copper + hydrogen oxide  
(water).

Similarly the oxide of iron obtained by heating iron in steam (page 62) is deprived of its oxygen, when it is heated in a current of hydrogen.

Iron oxide + hydrogen = iron + hydrogen oxide (water)  
The same chemical change may be considered either as reduction or as oxidation. If we consider the action of hydrogen on copper oxide, we may regard it either as the reduction of copper oxide by hydrogen or as the oxidation of hydrogen by copper oxide. The reduction of one substance proceeds simultaneously with the oxidation of some other. If our aim in an experiment is to obtain a metal or any other element from its oxide, the change is viewed as reduction; on the other hand, if our object is the formation of an oxide, the change is regarded as oxidation.

**2. Composition of water by mass.**—The electrolysis of water has shown that two volumes of hydrogen always combine with one volume of oxygen to form water. From the composition by volume we can easily deduce its composition by mass (weight), that is, the proportion by mass in which the two elements combine to form water. 1 litre of hydrogen (N.T.P.) has a mass of 0.09 grams; therefore the mass of two litres of hydrogen is 0.18 grams. 1 litre of oxygen (N.T.P.) has a mass of 1.429 grams. Now the ratio of 1.429 to 0.18 is 7.94 or 8 to 1 nearly. That is, 100 grams of pure water contain 88.8 grams of oxygen and 11.2 grams of hydrogen.

**Exp.**—This deduction as to the composition of water by mass (weight) is confirmed by reducing a known mass of copper oxide by the action of hydrogen on it, finding the mass of oxygen lost by the copper oxide, and collecting and weighing the water formed. A convenient form of apparatus for this purpose is shown in Fig. 42.

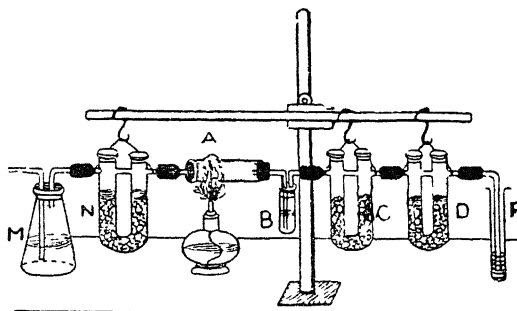


Fig. 42.

wire form, previously dried by ignition. B is a short test-tube in which most of the water collects. C and D contain dry calcium chloride to absorb any water that is not condensed in B. F is a test tube with a little concentrated sulphuric acid to show the rate at which hydrogen is passing through the apparatus. The tube A with its stoppers, connecting tubes, and the copper oxide, is carefully weighed; so are also B, C and D weighed together.

The parts of the apparatus are connected and a current of hydrogen from a generator is passed through, the rate being so regulated that about four bubbles pass out at F per second. After a few minutes the gas issuing from F is collected in a test tube and tested for the presence of air. If the hydrogen burns calmly, A is gently heated from end to end, so as to avoid cracking, then the temperature is gradually raised till the reduction of copper oxide begins. The heating is continued for some minutes after all the oxide has been reduced and all signs of moisture have disappeared from the combustion tube A. The flame is now removed, but the hydrogen is passed through till the tube has completely cooled. Finally a current of air is passed through the apparatus to sweep off any hydrogen present. Again A is weighed, the loss in mass is the mass of the oxygen which has combined

M and N contain substances to purify and dry the hydrogen from a generator. A is a bulb-tube of hard glass containing copper oxide in grains or in

with the hydrogen. B. C and D are also weighed together again ; the gain in mass is equal to the mass of water produced. The results of an experiment are given below :

the mass of A with oxide before heating = 30.56 grams.

„ „ copper after heating = 29.5 „

„ oxygen lost by the copper oxide = 1.06 grams

„ absorption tubes after experiment = 33.23 „

„ „ „ before „ = 32.03 „

„ „ of water formed „ = 1.20 grams

Hence 1.20 grams of water contain 1.06 grams of oxygen and (1.20—1.06) or 0.14 gram of hydrogen. From these data we get that 100 grams of water contain 88 grams of oxygen and 12 grams of hydrogen. Precise experiments give 88.81 grams of oxygen and 11.19 grams of hydrogen for 100 grams of water.

**The law of constant composition.**—Water is broken up into its elements oxygen and hydrogen (**analysis**) and is also produced by the union of its elements (**synthesis**). By whatever method pure water is obtained, 9 grams of water always contain 8 grams of oxygen and 1 gram of hydrogen. This constancy of composition is true not only of water but of every true chemical compound. This law of constant composition is stated as follows

**The same compound, howsoever it may be made, always contains the same elements combined in the same proportion by mass weight).**

This is the essential difference between a chemical compound and a mechanical mixture. The constituents of the latter may be mixed in any proportion, while the components of a definite compound always combine in the same **fixed** proportion. A mixture of oxygen and hydrogen may be made by mixing the constituent gases in any proportion we may like, but to produce water, oxygen and hydrogen must be present in the ratio of 1:2 by volume, or 8:1 by mass.

Hydrogen combines with oxygen to form two different oxides. One of these is water, which contains 8 parts by weight of oxygen for 1 of hydrogen. The second compound contains 16 parts by weight of oxygen for 1 of hydrogen, and is known as **hydrogen peroxide**. This readily decomposes into water and oxygen. The oxygen at the moment of liberation from this compound is energetic in action and destroys colouring matters by oxidation ; hence hydrogen peroxide is used to bleach bone, feathers, fur, hair, ivory and silk. Its oxidising character enables it to be used as an antiseptic in surgery and as a preservative of foods and drinks.

---

## XV

### COMBUSTION.

1. **Combustion.**—You have seen that chemical combination is generally accompanied by the evolution of heat. In some cases the disengagement of heat is evident to our senses, while in others it is not so. For example, when phosphorus or iron burns in the air, the disengagement of heat manifests itself to our senses, but when phosphorus smoulders in the air, or when iron rusts in the air, there is heat evolved but so slowly that it becomes diffused and does not manifest itself to our senses. But in both cases the action is the same.

The chemical combination of two or more substances with self-evident disengagement of heat and light is called **combustion** or **burning**. Where there is chemical combination without heat and light, there is **slow combustion**.

The amount of heat evolved in a chemical combination is the same, whether the combination results in a combustion or takes place without an appreciable rise of temperature. When a definite mass of iron rusts in the air, the heat of combination is not perceptible, but the total amount of heat is the same as is given out when the same mass of iron burns vividly in oxygen. In the former case the heat evolved is rapidly lost by conduction, convection and radiation; in the latter case, on the other hand, the combustion takes place so rapidly that the heat evolved is concentrated around the burning body and raises the temperature of the body itself or the products of combustion sufficiently to make them glow.

In ordinary burning the substance burns in the air and it wholly or partly unites with the oxygen of the air, producing one or more oxides. But we have seen

that substances can also burn in the absence of air. For example, iron burns in sulphur with a glow. What is necessary for combustion, is the presence of at least two substances which can chemically unite with each other with evolution of heat and light.

When one substance burns in the atmosphere of another substance, as in the burning of phosphorus, magnesium, or hydrogen in the air, the first substance is spoken as the **combustible** and the second as the **supporter of combustion**. Such a distinction, though convenient, is arbitrary. For, just as it is possible to burn hydrogen surrounded by the air, it is equally possible to burn the air at a jet surrounded by hydrogen.

**Exp,---**To the wider end of a lamp chimney fit a cork with two holes. Through one of these holes pass a short tube and connect it with a hydrogen generator. Through the other hole pass a metal tube so that one of its ends reaches to the middle of the chimney. Support

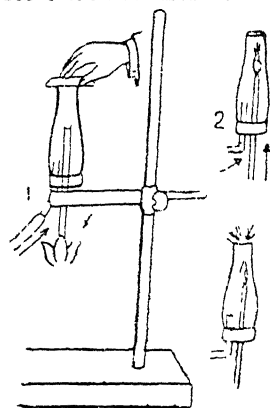


Fig. 43

the whole in a retort clamp, as in Fig. 43. By means of a piece of card-board lightly close the top of the chimney and pass hydrogen. The lighter hydrogen drives out the air and issues out of the lower end of the metal tube. Light the hydrogen at the lower end of the metal tube, when you are sure that all the air has been driven out. Here we have a flame of hydrogen burning in air (1). Now take away the card-board, the upward current of hydrogen draws the air through the metal tube and this carries the flame to the top of the metal tube and there we see a non-luminous flame of air burning in an atmosphere of hydrogen (2). The escaping gas may be lighted at the top of the chimney, and

we have simultaneously two flames, one of hydrogen burning in air at the top and another of air burning in hydrogen inside.

Whichever of the two combining substances may form the surrounding atmosphere, what is absolutely necessary for combustion is the presence of two or more substances capable of combining with one another.

**2. Flame.**—In ordinary burning we employ such materials as wood, coal, alcohol, petroleum, tallow, wax, &c. All these consist chiefly of hydrogen and carbon. If we hold a dry tumbler for a short time over the flame of any of these substances, it becomes bedewed with moisture. On pouring some clear lime water into the tumbler, the limewater becomes milky proving the presence of carbon dioxide. Combustible substances burn with or without flame. Substances which are gaseous or are rendered gaseous by the heat disengaged in burning, burn with a **flame**; all others merely **glow without a flame**. Hydrogen is gaseous at ordinary temperature; phosphorus, sulphur, magnesium, alcohol become gaseous during burning; wood, coal, tallow, oil evolve combustible gases due to their decomposition by the heat of combustion. On the other hand, iron, charcoal, coke yield no combustible gases and therefore merely glow. A flame is therefore a burning stream of gas, that is, a current of gas whose particles are continually combining with oxygen or some other surrounding gas with evolution of heat and light.

When two gases combine with evolution of heat and light, there is always a flame. The two gases may be made to combine with each other in two ways. The two gases may be mixed together in the requisite proportion, and the combination started by raising the temperature to the ignition point; the mixture then explodes with a flash. The explosion is due to the sudden expansion by heat or the production of a large quantity of gas within a small volume. Or, one gas may be made to stream out of a hole in a jet and burn



in an atmosphere of the other gas, as in a gas burner. In this case the flame will be steady, only when the gas takes fire as swiftly as it issues into the atmosphere. If the gas issues more swiftly, the flame goes out ; if it issues less swiftly, the flame strikes in.

The brightness of a flame is due to the presence of solid particles which become incandescent. Such is the case in mantle lamps. In the bunsen burner, the gas issues through a narrow jet into a wide tube, and just at the place where the jet opens, the wide tube has air-holes through which air is drawn in so that at the mouth of the wide tube burns an intimate mixture of gas and air. The flame is non-luminous. If we introduce into this non-luminous flame an incandescent mantle, the flame becomes luminous on account of the solids present in the mantle.

A luminous flame, like that of a candle, consists of three zones. The innermost consists of combustible gases ; the second one has incandescent solid particles in it and is therefore luminous ; the outermost is non-luminous, as complete combustion takes place in it.

---

## SULPHUR—SULPHIDES.

1. **Sulphur.**—Sulphur is, as you know, an odourless, tasteless, brittle solid of a pale yellow colour, insoluble in water. It is also a bad conductor of heat and electricity. When rubbed, it becomes electrified. It is a non-metallic element.

If sulphur is heated, it melts; on stronger heating it becomes dark brown and thickens. If the thick liquid is poured in a thin stream into cold water, it solidifies into a plastic mass, but becomes brittle again. At a still higher temperature sulphur boils and gives off a brownish yellow vapour. If this vapour is led into a cool vessel, it condenses to a very fine yellow powder, known as **flowers of sulphur**. All these are different physical forms of the same substance. (See page 12).

Sulphur occurs native in large quantities in volcanic regions mixed with earthy matter. Consequently sulphur has been known to man from antiquity. Native sulphur is heated in an oven and the sulphur vapour is led into large brick chambers, where it sublimes as flowers of sulphur. When the chambers have become sufficiently warm, the vapours condense to liquid sulphur, which is drawn off into moulds, in which it solidifies in sticks, known as **roll sulphur** (brimstone).

Though insoluble in water or alcohol sulphur dissolves easily in one of its own compounds, called **carbon disulphide**, an inflammable liquid formed by passing sulphur vapour over red hot charcoal. From its solution in carbon disulphide sulphur crystallises in octahedra (Fig. 44), different in form from the needle-shaped crystals obtained by fusion (page 11).



Fig. 44.

Sulphur burns in air or oxygen with a blue flame, forming a suffocating gas, called **sulphur dioxide** or **sulphurous anhydride** (page 51).

Sulphur + oxygen = sulphur dioxide.

Sulphur is used in the manufacture of sulphurous acid, sulphuric acid, sulphide of carbon. Large quantities of sulphur are consumed in the manufacture of matches, fireworks, gunpowder and the vulcanisation of rubber.

2. By heating sulphur or metallic sulphides in a current of hydrogen a gaseous compound of hydrogen and sulphur, **hydrogen sulphide**, with a very repellent smell, is formed. The same gas is produced in the putrefaction of animal and vegetable substances containing sulphur, such as eggs.

Hydrogen + sulphur = hydrogen sulphide (gas).

In a previous experiment (page 21) a compound was obtained by heating iron powder and flowers of sulphur. This compound is **iron sulphide**. On treating this compound with dilute sulphuric acid a foul-smelling gas was produced. This gas is the same **hydrogen sulphide**.

Iron sulphide + sulphuric acid = iron sulphate (green vitriol) + hydrogen sulphide.

Hydrogen sulphide burns in air forming water and sulphur dioxide, thus proving the presence of both hydrogen and sulphur in it.

Hydrogen sulphide + oxygen = water + sulphur dioxide.

In the laboratory hydrogen sulphide is used as a reagent and is prepared in an apparatus similar to Fig. 36 with iron sulphide instead of zinc.

Hydrogen sulphide turns blue litmus feebly red, and is therefore an acid. Hydrogen sulphide is soluble in water. The aqueous solution has the smell of the gas, and occurs in natural hepatic waters. From solutions of their compounds with other acids hydrogen

sulphide precipitates metals, such as silver, copper, lead, &c., as sulphides. Many of these are coloured.

Copper sulphate + hydrogen sulphide = copper sulphide (black) + sulphuric acid.

Metallic sulphides, when heated strongly in air, give off sulphur dioxide and become changed either into the metal or the oxide of the metal.

Mercury sulphide (cinnobar) + oxygen = mercury + sulphur dioxide.

Lead sulphide (galena) + oxygen = lead oxide + sulphur dioxide.

Many metals combine with sulphur to form sulphides. Copper foil burns in sulphur vapour forming black copper sulphide. Try.

3. Sulphides of many metals occur in nature and are known as **pyrites, glance, blende**. Pyrites and glances have a metallic lustre like galena (lead sulphide).

**Exp.**—Heat a piece of iron pyrites in a hard glass test-tube; in a short time flowers of sulphur sublime on the cold sides of the tube, proving the presence of sulphur. The residue is attracted by a magnet, which proves the presence of iron. Iron pyrites is a sulphide of iron.

---

## XVII

### SULPHUR DIOXIDE—SULPHURIC ACID.

1. **Sulphur dioxide.**—When sulphur is heated in an open vessel, it melts and catches fire, burning with a non-luminous blue flame, evolving a suffocating gas which causes coughing. This gas is, as you have seen, **sulphur dioxide**. Metallic sulphides, which largely occur in nature, burn in air giving off sulphur dioxide.

Sulphur dioxide is heavier than air. How can this be shown? Sulphur dioxide liquefies readily under pressure. This liquid vaporises at a very low temperature and therefore finds employment in ice-making machines.

A burning candle goes out when plunged into sulphur dioxide. As the gas itself is the result of combustion, it can neither burn nor support combustion.

It is largely soluble in water. How can this be shown? On account of its solubility in water it fumes strongly in moist air. Its solution is acid; it combines with water to form **sulphurous acid**. The gas and its solution destroy colouring matters and therefore find employment as **bleaching agents**, specially for wool, silk, paper, straw, and other delicate materials, such as are likely to be injured by other bleaching agents.

**Exp.**—Hold a rose over the gas issuing from burning sulphur. Observe the bleaching action of the gas.

The gas is used also as a disinfectant to destroy disease-producing micro-organisms.

Sulphurous anhydride combines readily with oxygen to form a higher oxide, namely, **sulphur trioxide** or **sulphuric anhydride**. In sulphur dioxide 32 grams of sulphur are combined with 32 grams of oxygen, where-

as in sulphuric anhydride, or sulphur trioxide, 32 grams of sulphur are combined with 48 grams of oxygen.

**Exp.**—Fill a jar with sulphurous anhydride by burning sulphur on a deflagrating spoon in a jar of oxygen. Why does the gas remain in the jar? Now pour some hot concentrated nitric acid into the jar. Reddish brown fumes of nitrogen peroxide are produced by the reducing action of sulphurous anhydride on nitric acid. Sulphur dioxide is therefore a **reducing agent**.

Nitric acid + sulphurous anhydride = sulphuric acid  
+ nitrogen peroxide.

**Exp.**—To some aqueous solution of sulphurous acid, add a few drops of barium chloride solution. The liquid remains clear. Leave the sulphurous acid solution exposed to air for some days and again test a portion of it with barium chloride solution. The liquid turns milky. The oxygen of the air has oxidised the sulphurous acid to sulphuric acid, which gives a white precipitate with a solution of barium chloride.

**2. Sulphuric acid.**—The ease with which sulphurous anhydride combines with oxygen to produce sulphuric anhydride forms the basis for the manufacture of sulphuric acid in large quantities.

**Exp.**—Set up the apparatus shown in Fig. 45. Through *a* and *b* are led sulphur dioxide and oxygen through concentrated sulphuric acid in *F*. The dry mixture of gases is then passed over heated platinised asbestos in *T*. The two gases combine and produce white fumes of sulphuric anhydride, which when dissolved in

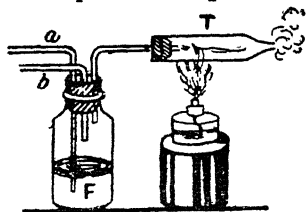


Fig. 45.

water form sulphuric acid. Finely divided platinum in the platinised asbestos acts as a catalyst.

Sulphurous anhydride + oxygen = sulphuric  
anhydride.

**Sulphuric anhydride + water = sulphuric acid.**

This method, known as the **contact process**, is employed in the manufacture of sulphuric acid. The older method of preparing sulphuric acid is based on the oxidation of sulphur dioxide by nitrogen oxides. Sulphur dioxide, steam, air, an oxide of nitrogen are brought together in a lead chamber, at the bottom of which collects a liquid rich in sulphuric acid, called *chamber acid*.

Sulphuric acid, or **oil of vitriol** of commerce, is a mixture of the compound **hydrogen sulphate**, which is the real sulphuric acid, with a more or less varying quantity of water. The most concentrated acid has the specific gravity 1.84. It is viscous and colourless, when pure.

Sulphuric acid mixes with water in all proportions, but the mixing must be effected with great care on account of the great heat evolved. The acid must always be poured in small quantities in water and stirred with a glass rod. Water should never be poured into the concentrated acid. A small quantity of water, coming in contact with the acid, is partially vaporised with a *sort of explosion which projects the hot liquid*.

Sulphuric acid is, even when cold, extremely corrosive and destroys animal and vegetable tissues by withdrawing the water which impregnates them.

**Exp.**—Pour some strong sulphuric acid on a splinter of wood. Observe how it becomes charred.

Dissolve 12 grams of sugar in 9 c.c. of warm water in a pretty large beaker. When the syrup has cooled, pour 12 c. c. of concentrated sulphuric acid rapidly into the beaker. The solution blackens, giving off at the same time large quantities of steam. The beaker becomes filled with a mass of porous carbon. Observe the inky smell.

Concentrated sulphuric acid absorbs aqueous vapour. Gases are dried by making them pass through

tubes filled with pieces of pumice stone impregnated with concentrated sulphuric acid. Solids, or aqueous solutions, which we desire to dry or evaporate, are placed under a dessicator, which is a bell jar containing a dish of the concentrated acid.

**Exp.**—Place a few crystals of borax in a dessicator and allow them to stand for a day or two. You will notice that the crystals effloresce owing to the loss of their water of crystallisation, as the air is rendered dry by the concentrated sulphuric acid in the dessicator.

Sulphuric acid is decomposed, when very strongly heated, into oxygen, water and sulphurous anhydride.

Hydrogen sulphate=sulphurous anhydride+oxygen  
+water.

Sulphuric acid deeply reddens blue litmus, thus showing its strongly acid character.

Among chemical products sulphuric acid finds the most numerous and most important applications. The energy with which it combines with metals renders it capable of decomposing their compounds with the other acids; it combines with the metal to produce a sulphate, and the acid of the compound attacked is liberated; hence it is constantly used for the preparation of other acids. You will come across some examples in the sequel. Sulphuric acid serves to prepare sulphates, such as those of sodium and ammonium, which enter into the composition of chemical manures. It serves as the active liquid in many voltaic batteries.

**3. Sulphates.**—Sulphuric acid attacks a large number of substances, especially metals. You have seen how hydrogen is produced by the action of dilute sulphuric acid on some easily oxidisable metals, such as zinc, iron, or magnesium. The compounds formed are called **sulphates**.

Zinc+hydrogen sulphate=hydrogen+zinc sulphate.



**Exp.**—Add magnesium little by little to dilute sulphuric acid till the evolution of gas ceases, and then evaporate the solution. Crystals of magnesium sulphate separate from the solution. This salt is known in commerce as **Epsom salts** and is generally prepared by the action of dilute sulphuric acid on magnesium carbonate (dolomite, magnesite) which occurs in nature. Epsom salt finds employment in medicine.

Concentrated sulphuric acid when heated with metals evolves sulphur dioxide with formation of the metallic sulphate.

Copper + sulphuric acid = sulphuric dioxide + water +  
copper sulphate.

**Exp.**—Heat some copper turnings with concentrated sulphuric acid. Observe the fumes of sulphur dioxide from the smell. Observe the dark-brown colour of the solid residue. When cool, pour the residue into a beaker of water, and allow it to stand. Crystals of blue vitriol are formed, as the liquid evaporates. Why was not the solid residue left after heating blue?

Gold and platinum are the only metals not attacked by sulphuric acid.

Sulphuric acid acts also on the oxides of metals producing sulphates.

**Exp.**—Put some black oxide of copper in a beaker containing a little water. It does not dissolve. Then add sulphuric acid; the oxide disappears little by little and the liquid turns blue. On heating the oxide dissolves more rapidly. The blue liquid contains **sulphate of copper** (blue vitriol).

You may repeat the experiment with oxide of zinc (**zinc white**), and obtain crystals of **zinc sulphate** (white vitriol).

Copper oxide + sulphuric acid = Copper sulphate +  
water.

Zinc oxide + sulphuric acid = Zinc sulphate + water.

Sulphuric acid gives with a solution of barium chloride a white precipitate of barium sulphate, which is insoluble in water and acids.

Hydrogen sulphate+barium chloride=barium sulphate+hydrogen chloride.

Aqueous solutions of sulphates also give the same precipitate with a solution of barium chloride.

Green vitriol and blue vitriol are two well-known coloured sulphates. The former is the sulphate of iron (ferrous sulphate), the latter that of copper. Both are soluble in water and can be obtained as crystals by the slow evaporation of their aqueous solutions. Other important sulphates are **Glauber's salts** and **gypsum**. The former finds employment in medicine. The latter by heating loses water of crystallisation and becomes the powdery **plaster of Paris**, which is capable taking back its water of crystallisation and becoming hard. The former is sodium sulphate and the latter is sulphate of calcium.

---

## XVIII

### CARBON.

1. **Carbon.**—In studying the properties of oxygen you saw that a piece of charcoal burning in oxygen produces a gas which turns lime water milky (page 50). Similarly in studying the burning of a candle in air and in oxygen you saw the material of the candle yield by its combustion the same gas (page 15). This property is possessed by all substances which contain an element, called **carbon**.

There are different varieties of carbon, but all of them yield on burning in air or oxygen carbonic acid gas which has the property of turning lime water milky.

Some of these, such as diamond, graphite (or plumbago), coal, peat, occur in nature; others, such as wood-charcoal, bone-black, lamp black, gas carbon, coke, sugar carbon, are artificially made.

**Diamond** is almost pure carbon. It is a clear transparent stone with a crystalline form [octohedron]. It scratches all other substances, and is employed for cutting glass and polishing hard materials. On account of its shining quality when cut it is used in jewelry (**brilliants**). When heated to redness in a current of oxygen, diamond disappears slowly producing carbonic acid gas and a very minute quantity of ash.

**Graphite** or **plumbago** (blacklead) is a soft, greyish black, lustrous solid, slimy to the touch. It leaves a black streak when drawn on paper, and hence employed in making black-lead pencils. It is used also as a lubricant to diminish friction, as a coating for cast iron articles to prevent them from rusting, and in the manufacture of crucibles for high temperatures. It is less pure than diamond, leaving always in burning a

residue of ash. It is often crystalline. It is a conductor of heat and electricity.

Plants and their materials, buried under ground for millions of years, have under the influence of the earth's heat, undergone a slow decomposition and have become transformed into **coal**, which is composed mainly of carbon.

**Exp.**—Fit up apparatus as in Fig. 46. Put in the dry test-tube some fragments of coal, and heat it strongly. A smoky gas issues from the jet after some minutes. Apply a light to this. It takes fire and burns with a luminous flame. This gas, when purified, is known as **coal gas**. When no more inflammable gas comes off, withdraw the flame. After the tube has cooled, examine its contents. The residue is a greyish-black porous mass, resembling charcoal. This is the artificial variety of carbon, known as **coke**.

Coke contains all the incombustible ash of coal and about 91 per cent of carbon. It is a very good combustible, generating much heat. While coke is found at the bottom of the iron retort in which coal is heated, at its top is formed a hard stone-like deposit, called **gas carbon**. This contains about 95 per cent of carbon. As it has the power of conducting electricity, it is used in electric batteries and for the carbon rods of arc-lamps.

**Exp.**—Fit up apparatus as in Fig. 46. Put a few pieces of wood in the test-tube and heat strongly. The wood suffers **destructive distillation**. An oily liquid collects in the cylinder, and a gas escapes through the jet, which, when lighted, burns with a bright flame. There is left in the test-tube a black residue, which is pure **wood charcoal**.

On a large scale wood charcoal is prepared either by the incomplete combustion of wood itself in a

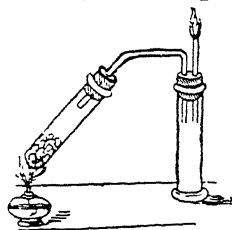


Fig. 46.

pit or kiln with access of a little air, or by heating wood in closed retorts without access of air.

Wood charcoal is not pure carbon; on burning it leaves ash (potash). It has the fibrous structure of wood and is very porous. It has the remarkable property of absorbing some gases and hence used as a disinfectant.

**Exp.**—Heat a piece of bone or leather in a test-tube (Fig. 46). The substance chars, showing the presence of carbon, and an oily liquid again collects. The black residue in the test-tube forms **animal charcoal** or **bone-black**.

Bone-black contains only about 12 per cent of carbon, but this is in a very finely divided condition spread over the mineral matter (bone ash) contained in the bone. It is used in industry as a decolorising agent, *e.g.*, in removing the colouring matter from jaggery extracted from the sugar-cane.

**Exp**—Put 3 or 4 grams of bone-black into a flask, and pour over it 10 or 15 c.c. of a solution of blue litmus (indigo, magenta). Shake the flask well and boil. Then filter. Observe that the filtrate is colorless.

**Exp.**—Pour a little kerosin oil in a tin dish and light it. It burns with a smoky flame, the smoke being due to the presence of unburnt particles of carbon. Hold a porcelain dish or a glass plate over the flame. A black deposit is formed. This deposit consists of **lamp-black**. In industry it is prepared by the incomplete combustion of oils and resins in large brick chambers. All the hydrogen and a portion of the carbon are burnt, but the rest of the carbon collects as soot, which is purified by heating strongly without air.

Lamp-black is carbon in fine powder, and is employed in manufacturing printer's ink and black pigments.

When sugar is heated in a test-tube, inflammable gases are evolved and the residue is **sugar carbon**, which

contains no impurities ; sugar is a compound of carbon, hydrogen and oxygen, and can be obtained perfectly pure. Its complete decomposition leaves behind only pure carbon. Sugar carbon must be treated as a type of wood charcoal, for as you know the sugar of commerce is obtained from plants (sugar-cane, beet root, palmyra).

**2. Properties of carbon.**—All forms of carbon are infusible. They are all combustible. 12 grams of diamond or sugar carbon, burning in oxygen, yield 44 grams of carbonic acid gas. The combustion of carbon generates a large amount of heat. The heat obtained by burning 1 kilogram of carbon will raise the temperature of 80 kilograms of water from 0° to 100° C. The principal use of carbon is therefore the production of heat by burning.

On account of the readiness with which carbon combines with oxygen, carbon reduces a great number of oxygen compounds.

It reduces water. This can be seen by arranging apparatus as in Fig. 32 for the decomposition of water by iron. This action is utilised in industry for producing a combustible gas (called **water-gas**) by passing steam mixed with air over red-hot carbon. The product is a mixture of hydrogen, carbonic acid gas and another gas (carbon monoxide), which is inflammable.

Why does a small quantity of water thrown on a glowing fire increase the combustion ?

Many metallic oxides are reduced by carbon. In experiments on pages 71, 72, you reduced the oxides of tin and copper by means of carbon.

Tin oxide + carbon = carbon dioxide + tin.

Copper oxide + carbon = carbon dioxide + copper.

The reduction is not always so easy. For example, the oxide of zinc has to be heated much more strongly in earthen-ware apparatus and carbon monoxide is evolved.

It is by reduction with carbon of the naturally occurring oxides or the oxides obtained by roasting the sulphides that some metals (iron, zinc, tin) are industrially prepared.

Certain oxides, such as lime, can be reduced only at the very high temperature of the electric furnace. But then the metal set free combines often with the carbon to form a **carbide**, such as the **carbide of calcium** used for generating **acetylene**, an inflammable gaseous compound of carbon and hydrogen.

Carbon (coke) reduces also sand, an oxide of silicon, in the electric furnace by forming carbon dioxide and **carborundum** (silicon carbide). Carborundum is extremely hard and is used as a substitute for corundum or emery.

On heating wood charcoal with concentrated sulphuric acid the latter is reduced to sulphur dioxide. Similarly nitric acid is reduced by carbon.

Carbon + sulphuric acid = carbon dioxide + sulphur dioxide + water.

Carbon reduces carbon dioxide itself. When a current of dry carbonic acid gas is passed over red-hot carbon, it loses oxygen and becomes reduced to **carbon monoxide**.

Carbon + carbon dioxide = carbon monoxide.

In carbon dioxide 12 grams of carbon are combined with 32 grams of oxygen, whereas in carbon monoxide the same mass of carbon is combined with 16 grams of oxygen. Carbon monoxide is a combustible gas, which burns in air with a blue flame, forming carbon dioxide. The blue flame at the top of a coal fire is due to this gas which is produced by the carbon dioxide formed at the lower part of the fire being reduced in passing through the upper layers of red-hot carbon. Carbon monoxide is insoluble in water, has no acid properties and is poisonous.

Carbon forms many compounds with hydrogen. One of them is acetylene, obtained by the action of

water on calcium carbide, and used as an illuminant in cycle lamps.

Calcium carbide + water = slaked lime + acetylene.

Oxy-acetylene flame is very hot and is used for cutting steel plates.

Another compound of carbon and hydrogen is **marshgas** (methane) which is seen bubbling up from marshes. This is the cause of explosion in coal mines.

If sulphur vapour is passed over carbon heated to redness, gaseous sulphide of carbon is formed, which condenses in water to a liquid which does not dissolve in water. **Carbon disulphide** is, as you already know, a colourless liquid of a disagreeable odour, capable of dissolving sulphur, phosphorous, rubber, fatty substances. Carbon disulphide, liquid as well as vapour, is highly inflammable. Its combustion yields carbonic acid gas and sulphurous anhydride and in the insufficiency of oxygen also sulphur.

---



## XIX

### CARBONIC ANHYDRIDE.

You have already seen that when charcoal is burnt in air or oxygen, a gas is produced, which dissolves to a slight extent in water, forming a feeble acid, and turns lime water milky. The same is the case with every form of carbon, or any organic compound of carbon. The gas is, as you know, a compound of the elements, oxygen and carbon, called **carbon dioxide** or **carbonic anhydride**.

**Exp.**—Pass oxygen through a hard glass combustion tube containing a piece of charcoal and heat the charcoal. The charcoal glows and becomes consumed. Collect the gas produced by downward displacement in a tall jar. Apply tests.

Carbon dioxide is a colourless, odourless gas. It is heavier than air and can therefore be kept for a long time in open jars. It can be poured from one jar into another, as if it were a liquid.

A burning candle goes out at once in carbon dioxide. Carbon dioxide turns lime water milky. The milkiness is due to the formation of a compound of carbon dioxide and the lime in lime water. This compound is calcium carbonate (chalk), which is insoluble in water, and is therefore thrown down as a white precipitate.

**Exp.**—Place a lighted candle on the table. Pour the carbon dioxide from a jar over it; the flame is extinguished by the fall of carbonic anhydride on it. This property of going down on account of carbon dioxide being heavier than air is taken advantage of in extinguishing flames.

Water dissolves its volume of carbon dioxide. Soda water and other aerated waters contain five to six

times their volume of carbon dioxide dissolved **under pressure**.

Carbon dioxide is easily liquefied by pressure and liquid carbon dioxide is sold in steel cylinders. Aerated waters are made with the gas from these cylinders. The gas for these cylinders is obtained by combustion or by the fermentation of solutions of sugar (molasses) with yeast. By this fermentation alcohol and carbon dioxide are produced. **Alcohol** (known also as **spirits of wine**) is a colourless, inodorous, mobile liquid, lighter than water, which forms a good solvent for many substances (such as shellac, camphor, iodine) which are insoluble in water. It is a compound of hydrogen, carbon and oxygen.

The aqueous solution of carbon dioxide turns blue litmus feebly red. It is acid. The aqueous solution may be regarded as containing carbonic acid, the real acid compound of carbon, oxygen and hydrogen.

Water charged with carbonic anhydride dissolves calcium carbonate (chalk).

**Exp.**—To a beaker containing clear lime water add a few drops of soda water. A turbidity is produced, which disappears on adding more of the soda water. The first drops of soda water produce a precipitate of calcium carbonate, as it is insoluble in water, but on adding more the excess of carbon dioxide dissolves the precipitate.

Boil the clear liquid thus obtained. The turbidity reappears, as the excess of carbon dioxide is expelled by boiling.

As you have already seen (page 56), water containing calcium carbonate in solution is **hard**, but as its hardness can be removed by boiling, it is spoken of as **temporary hardness** to distinguish it from the hardness of water containing calcium sulphate (gypsum) in solution, which cannot be removed by boiling and is therefore called **permanemt hardness**.

Solutions of alkalis (caustic soda, caustic potash) readily absorb carbon dioxide forming carbonates (potash, soda).

The carbonates of all metals (chalk, marble, soda, potash) are easily decomposed by dilute acids (sulphuric acid, hydrochloric acid).

Metal carbonate + hydrogen chloride = metal chloride + hydrogen carbonate.

Hydrogen carbonate, that is, carbonic acid, decomposes rapidly into carbon dioxide and water.

When potassium is heated in a current of carbon dioxide, potassium burns liberating carbon. The potassium oxide formed absorbs some of the carbon dioxide and forms potassium carbonate. By dissolving the residue in water, we obtain black particles of carbon.

When carbon dioxide is passed over red hot iron in an apparatus similar to that employed in preparing hydrogen from steam and iron, carbon monoxide is produced with formation of oxide of iron.

In the laboratory carbon dioxide is prepared generally by the action of dilute hydrochloric acid on chalk or marble, when a brisk effervescence takes place due to the evolution of carbon dioxide :

Calcium carbonate + hydrogen chloride = calcium chloride + carbon dioxide + water.

The apparatus is the same as that employed for the preparation of hydrogen (page 63). An ordinary two-necked Woulff's bottle may be used for a flask with a stopper with two holes. As the gas carries with it traces of hydrochloric acid, it should be washed by bubbling it through water in a wash-bottle. The gas being heavier than air can be collected by downward displacement.

---

## XX

### CARBON DIOXIDE IN THE AIR.

1. Besides nitrogen and oxygen there is always present in the air a certain amount of carbon dioxide. There are about 3 volumes of carbon dioxide in 10,000 volumes of air.

**Exp.**—Pour some clear lime water in a shallow dish and leave it exposed to the air for some hours. Observe that, while the lime water in the bottle remains clear, the liquid in the dish has become turbid through carbon dioxide.

This carbon dioxide must necessarily have come from the air. We can easily account for the presence of carbon dioxide in the air. When any substance containing carbon burns in the air, carbon dioxide is emitted into the air. In all common combustions the combustibles (wood, charcoal, coal, petroleum, alcohol, coal gas, water gas) contain carbon. In fermentation and putrefaction much carbon dioxide is produced. Carbon dioxide issues from fissures in the ground in volcanic regions.

2. **Respiration.**—It is a matter of ordinary experience that men and animals can not live without a constant supply of air any more than a candle can burn without fresh air. Just as a burning candle soon goes out in a confined volume of air, so man and other animals die in a closed space without free access of air. This dependence on fresh air, common to a burning candle for its continuous burning and to an animal for continuous life, indicates some similarity between the processes of burning and respiration.

**Exp.**—To see whether the air in breathing undergoes the same change as when a candle burns in it, perform an experiment. Fit up a test-tube with cork and tubes, as in Fig. 47, and put in it some clear lime water. Suck at the end of the shorter tube so that fresh air bubbles through the lime water. Observe that the lime water does not turn milky. Blow air from the lungs through the longer tube. Observe that the lime water becomes turbid. It therefore follows that men as well as animals breathe in fresh air and breathe out carbon dioxide into the air.

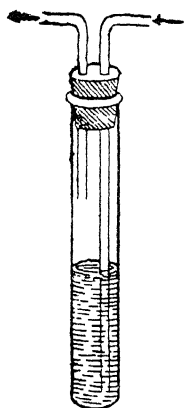


Fig 47.

The respiration of animals and plants brings about a veritable slow combustion in their organs. It is the heat produced by this slow combustion that maintains our bodies at a constant temperature. The food we take suffers in our organs profound changes; a part of the carbon and the hydrogen they contain is burnt. The oxygen of the fresh air we breathe in comes in contact with the blood in the lungs. The blood absorbs this oxygen and becomes bright red. This oxygenated blood is carried by circulation to all the organs where the oxygen is taken up by the tissues and the waste products of their activity, such as water, carbon dioxide and ammonia, are carried away by the blood. In the course of the blood through the lungs water and carbon dioxide are removed in the exhaled air, while the inhaled air supplies fresh oxygen to the blood. Expired air contains 4.03% by volume of carbon dioxide, 17% of oxygen, and 78% of nitrogen.

Despite the large quantity of carbon dioxide sent into the air by combustion, respiration, fermentation, putrefaction, volcanic action, the air contains only 0.03% of carbon dioxide. How is the excess of carbon dioxide removed? It is a well-known fact that plants contain

carbon; any part of a plant when heated chars. Do plants get the carbon necessary for their growth from the ground? It cannot be, for there are many plants that grow in water. Do they then obtain it from the carbon dioxide contained in the air? The following experiment answers the question in the affirmative.

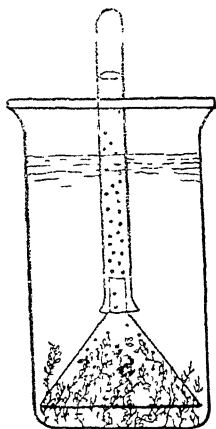


Fig. 48.

**Exp.**—Fill a large funnel with some fresh water-cress (*nasturtium officinale*) or pondweed (*elodea canadensis*) and slide it down into a large beaker of water saturated with carbonic acid gas. By tapping drive off any air held between the leaves. Place a test-tube full of water over the stem of the funnel (Fig. 48). Expose the whole to sun-light. Observe the gas accumulating in the test-tube, and test it with a red-hot splinter. The rekindling of the splinter proves the gas to be oxygen.

**Exp.**—Put plenty of green leaves in a trough containing water and cover it with a tubulated bell-jar. Attach a wire carrying a taper to a stopper fitting the mouth of the bell jar. Light the taper and insert it into the jar pressing the stopper tight. After some time the taper goes out. Why? The taper is taken out, relighted, and again inserted into the jar, pressing the stopper tightly in as before. The taper goes out at once. Why? Now set aside the bell-jar in sun-light for some hours. Then remove the stopper and introduce the lighted taper into the bell-jar. The taper continues to burn. Why?

All green plants in the presence of sun-light decompose carbon dioxide into carbon and oxygen. They retain the carbon for their growth and return the oxygen to the air. Man and animals inhale

oxygen from the air; this oxygen combines with the carbon in their bodies and is given back to the air as carbon dioxide. Thus a circulation of carbon is kept up in nature.

---

## XXI.

### CHALK—POTASH—SODA—CARBONATES.

1. **Chalk.**—As we have so often seen, calcium carbonate is formed by the action of carbon dioxide on lime water.

Chalk, marble, lime-stone, shells are natural varieties of calcium carbonate. By strongly heating any of these we get quicklime, which is an oxide of calcium, and carbon dioxide escapes. By burning large quantities of shells and lime-stone in kilns quicklime is prepared.

**Exp.**—To show that carbon dioxide is given off when chalk or any other variety of calcium carbonate is heated strongly, put some pieces of chalk into the bowl of a clay pipe and close the mouth of the bowl well with moistened clay. On drying heat the bowl in a furnace to bright redness. By means of a rubber tube connect a glass tube to the narrow end of the claypipe and lead the gas evolved into lime water. The lime water turns milky, showing that the gas given off is carbon dioxide. Hence chalk or marble is a compound of quicklime and carbon dioxide.

Chalk=quicklime + carbon dioxide.

As quicklime is a compound of oxygen and the metal calcium, and as carbon dioxide is a compound of carbon and oxygen, chalk (marble, lime-stone) is a compound of calcium, carbon and oxygen.

2. **Potash.**—For washing purposes a solution of a white substance, called **potash**, is often employed. Potash dissolves very easily in cold as well as warm water. The solution is slimy to the touch and affects the skin.

**Exp.**—Introduce a small quantity of this substance on a platinum wire or asbestos thread into a non-



luminous flame. The flame is coloured violet, reminding you of the colour of burning potassium. Potash therefore contains potassium.

**Exp.**—Put some potash in a jar and pour over it dilute sulphuric acid or hydrochloric acid. There is effervescence, indicating the evolution of gas. Plunge a burning candle into the jar; it goes out at once, showing the presence of carbon dioxide. Lime water turns milky in the gas.

**Potash is therefore a carbonate of potassium.**

Potash is obtained by dissolving wood ashes in hot water and evaporating the solution. At the present day potassium carbonate is prepared mostly from the large quantities of naturally occurring potassium chloride by the action of sodium carbonate on it. Crystals of potassium carbonate are deliquescent.

**3. Soda.**—Soda, or washing soda, is used more frequently than potash for washing purposes. It is sold in the form of colourless, transparent crystals. On exposure to the air the crystals lose their transparency, their surface becoming covered with a white, opaque layer. This efflorescence is due to the crystals losing their water of crystallisation.

Soda is easily soluble in warm water; the solution has the taste of lye and has a corrosive action on the skin.

Soda dissolves fats and is therefore capable of cleansing greased goods.

On the addition of soda to water the water becomes soft, that is, a lather is easily produced, which has a cleansing effect.

**Exp.**—Introduce a little washing soda on an asbestos thread into a non-luminous flame. The flame becomes golden-yellow, a colouration which is characteristic of sodium and its compounds. Washing soda therefore contains sodium.

**Exp.**—Place some soda crystals in a gas jar and pour over it dilute sulphuric acid. There is effervescence. A burning candle introduced into the jar goes out, showing that the gas given out is carbon dioxide. Also lime water turns milky in the gas.

Hence **washing soda is a carbonate of sodium.**

Sodium carbonate is manufactured on a large scale from common salt, and this accounts for its cheapness.

**Exp.**—Pass a stream of carbon dioxide through a solution of washing soda for some time. A precipitate is thrown down in the form of a white powder. This is a second carbonate of sodium, which contains for the same mass of sodium double the mass of carbon dioxide. Hence it is called **bicarbonate of sodium.**

While washing soda is highly alkaline, the bicarbonate is slightly alkaline, and contains hydrogen. This is known as **baking soda**, as it forms the chief ingredient of baking powders. Baking soda, on heating, gives up a part of its carbon dioxide and this causes the dough to swell as in "pappadums."

Formerly the bicarbonate was used in the making of aerated water. To make the carbonate give up its carbon dioxide, tartaric acid is added. A mixture of baking soda with powdered tartaric acid forms what is known as **effervescent powders**. From the fact that formerly aerated water was prepared from baking soda, the ordinary aerated water which contains no soda but only carbon dioxide is spoken of as **soda water**.

When a concentrated solution of washing soda is mixed with slaked lime, it becomes converted into **caustic soda**; the lime is precipitated as carbonate and collects on the bottom; the decanted liquid contains the caustic soda. This liquid, concentrated and treated with the proper quantity of cocoanut or palm oil, yields soap, which is used for toilet and washing purposes.

In carbon-dioxide fire extinguishers a solution of sodium bicarbonate is employed. The apparatus contains a bottle of sulphuric acid attached to its top and a solution of baking soda at its bottom. On inverting

the extinguisher or by striking some part of it the acid is dropped into the solution; the carbon dioxide evolved forces out through a nozzle a solution which extinguishes the fire better than ordinary water.

Most metallic carbonates, when heated strongly, become converted into oxides. For example, lead carbonate (white lead) on heating turns into the yellow oxide, litharge.

All carbonates are decomposed by dilute acids with evolution of carbon dioxide and the formation of the corresponding salt of the acid.

Most carbonates are insoluble in water.

---

## NITROGEN—AMMONIA.

1. **Nitrogen.**—Nitrogen and oxygen are the principal constituents of the air. there are always present in air but the some carbonic acid gas and aqueous vapour. As you have already seen, nitrogen exists in the air in the free state. We separate it from the air by passing a slow stream of dry air over heated copper. This nitrogen differs in density from nitrogen prepared from some of its compounds. It has been shewn that nitrogen separated from the air contains very small quantities of certain rare gases (argon, krypton, &c.).

Nitrogen is a colourless, odourless, tasteless gas, slightly lighter than air, neither inflammable nor capable of supporting combustion. It is chemically very inactive at ordinary temperatures, but at high temperatures, as that of the electric spark or electric arc, it combines with oxygen forming reddish-brown peroxide of nitrogen.

Nitrogen present in the air appears to act as a moderator of the combustions in air. This nitrogen seems essential for life, as in pure oxygen animals live very rapidly and die.

Nitrogen is an indispensable component of the tissues of plants and animals. All nitrogenous substances excreted by living beings or remaining in their bodies after death are transformed into less complex substances, or even into nitrogen under the action of microbes. One of these products is **ammonia**, a pungent gas which turns red litmus paper blue. This is the only alkaline gas known.

**Exp.**—Heat gently in a test-tube some hair, horn, or leather with some caustic soda, and place a strip of red litmus paper on the mouth of the tube. The paper turns blue, showing that an alkaline gas is evolved. The alkaline gas is ammonia.

Ammoniacal products are converted by other microbes into nitric acid and nitrates.

2. **Ammonia.**—The gases coming from the retorts in which coal is subjected to destructive distillation for producing coal gas are passed into water, which absorbs some of the gases. This impure gas liquor is the chief source of ammonia for commercial purposes.

**Exp.**—Heat some coal strongly in a hard glass test-tube. Test the tarry liquid which is deposited with red litmus paper. Note its alkaline character. This liquid contains ammonia dissolved in it.

To remove impurities from the gas liquor it is mixed with lime and heated. Ammonia gas is evolved and is dissolved in water. This solution is commercially known as **liquor ammonia**. Instead of dissolving the ammonia in water the gas is often passed into tanks containing hydrochloric acid or sulphuric acid. In the former case **sal-ammoniac** or ammonium chloride is produced, and in the latter ammonium sulphate is formed. Sal-ammoniac is used in batteries in telephone and electric bell circuits, while ammonium sulphate is employed as manure.

**Exp.**—Examine some sal-ammoniac from the bazaar. Note that it is crystalline and inodorous. Heat a small quantity in a dry test tube; it volatilizes and forms a sublimate. Put a small quantity of it in water; it readily dissolves. Test the solution with blue and red litmus paper; the solution is neutral. Add to a small portion of the clear solution a few drops of a solution of silver nitrate; the immediate formation of a thick curdy white precipitate, slowly

changing colour to black, indicates the presence of a chloride.

**Exp.**—Mix some dry quicklime with about half its volume of powdered sal-ammoniac and heat the mixture in a test-tube. There is no sublimate, but ammonia gas is evolved, which is recognised at once by its odour and its action on red litmus paper.

Hence sal-ammoniac is a compound of ammonia and hydrogen chloride. It is called ammonium chloride, by analogy with sodium chloride, which is formed by the action of hydrochloric acid on caustic soda solution.

Ammonium chloride + quick lime = calcium chloride  
+ water + ammonia.

**Exp.**—Produce some ammonia gas as in the last experiment keeping the heated tube vertical. Invert over its mouth a wider tube containing a strip of moist red litmus paper. The paper turns blue showing that ammonia is lighter than air and can be collected by upward displacement.

Collect the gas in another tube in a similar manner and lifting the tube insert its mouth in water. The water rushes in with violence, showing the great solubility of ammonia in water. This experiment may be arranged in the form of a fountain as shown in Fig. 49. The bottle contains ammonia gas and the tumbler contains red litmus solution.

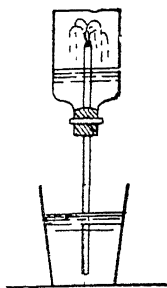
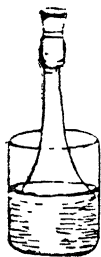


Fig. 49.

A solution of the gas in water may be obtained by attaching the delivery tube by rubber tubing to a funnel and placing the latter in such a way that its mouth just dips under the surface of water (Fig 50). This will prevent the water from rushing back into the generating flask, because when the water rises a little in the funnel the level of water



outside falls below the mouth of the funnel and the air enters

Ammonia is a colourless gas with a very pungent smell and is highly soluble in water; the aqueous solution is alkaline.

The aqueous solution of ammonia smells strongly of ammonia. On warming the solution it easily gives up the dissolved ammonia.

Ammonia liquefies on cooling and compressing it. Liquid ammonia is employed in the manufacture of ice to produce cold by its

Fig 50. evaporation.

If dry ammonia gas is passed through a heated combustion tube containing black oxide of copper, then through an empty wash-bottle, and finally received in an inverted jar of water in a pneumatic trough, water collects in the wash-bottle and nitrogen collects in the jar, thus proving that ammonia contains hydrogen and nitrogen. The oxide of copper becomes reduced to metallic copper.

The aqueous solution of ammonia combines with various acids to form the corresponding salts of ammonium.

A well-known compound of ammonia is **sal-volatile** or ammonium carbonate, which is used in smelling salts.

---

## XXIII

### COMMON SALT. SODIUM.

1. **Common salt.**—In every household is found the indispensable common salt, used for flavouring foods and for pickling purposes. Domestic animals also are eager after common salt.

Common salt is a white solid, sold either as a fine powder (table salt) or in large grains (cooking salt), which is easily soluble in water.

On evaporating a saturated solution of common salt it crystallizes out in the form of cubes without any water of crystallisation. On heating in a spoon some large crystals splitting sounds are produced and some small pieces are thrown out. This decrepitation is due to the evaporation of water enclosed in the crystals during crystallisation. Common salt is difficult to melt.

Common salt occurs very abundantly in nature, in the solid state as **rock salt**, which is dug from mines, and in the dissolved state in the oceans and seas. A cubic metre of seawater contains from 26 to 31 kilograms of common salt. From sea water the salt is obtained by evaporation in salt pans.

Impure common salt contains traces of magnesium and calcium chlorides, which cause the absorption of moisture from the air.

**Exp.**—Introduce a little common salt on asbestos thread into a non-luminous flame. The golden yellow colour of the flame indicates the presence of sodium.

2. **Sodium.**—Sodium is prepared on a large scale by the decomposition of fused common salt by a strong electric current with carbon electrodes. Metallic



sodium is liberated at the negative pole, while a greenish yellow gas, **chlorine**, is liberated at the positive pole. This chlorine is an element, forming the other component of common salt.

The chlorine produced at the positive electrode is collected under pressure in steel cylinders and sold in the market.

An aqueous solution of sodium chloride is similarly decomposed by the electric current. But the sodium liberated at the negative electrode at once combines with water forming **caustic soda** and sets free hydrogen. Chlorine collects at the positive electrode, if the electrodes are isolated. If they are not isolated, the chlorine spreads through the liquid, combines with the caustic soda forming a yellow bleaching liquid. If the solution is hot, the action of chlorine on caustic soda produces sodium chlorate, a substance quite resembling potassium chlorate in properties.

The solution of caustic soda obtained by the electrolysis of a cold solution of sodium chloride is evaporated to dryness, melted and moulded into sticks. **Caustic soda** is a white solid, slimy to the touch and corrosive; it dissolves in water with evolution of heat. It is highly alkaline.

By the action of ammonia and carbon dioxide on a solution of common salt under pressure, baking soda (sodic hydric carbonate) is precipitated. By heating the latter to redness anhydrous sodium carbonate is obtained. Ammonia changes into ammonium chloride.

Potassium chloride behaves exactly like sodium chloride. At the present day large quantities of potassium chlorate and caustic potash are manufactured by the electrolysis of potassium chloride. As potassium chlorate is more soluble at a higher temperature, it crystallises out on cooling.

---

## XXIV.

## HYDROGEN CHLORIDE.

Sulphuric acid acts vigorously on common salt.

**Exp.**—Put some common salt in a test tube and pour a few drops of sulphuric acid on it. Observe the brisk effervescence. The gas evolved has a pricking odour and fumes strongly in the air. The fumes redden blue litmus paper. Remember that sulphuric acid does not give off fumes. This gas is **hydrochloric acid gas**, whose chemical name is **hydrogen chloride**.

Sodium chloride + hydrogen sulphate = sodic hydric sulphate + hydrogen chloride.

Hydrogen chloride is prepared in the laboratory by this reaction.

**Exp.**—Fit up the apparatus shown in Fig. 51.

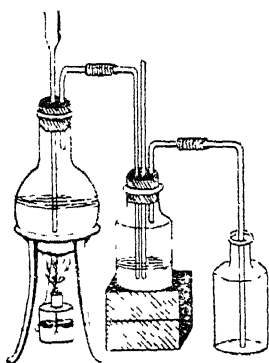


Fig. 51.

gas before it is collected.

Put into the flask 50 grams of common salt and cover it with a solution of 90 grams of sulphuric acid in 20 grams of water. Apply gentle heat. Hydrogen chloride being heavier than air can be collected by downward displacement. The rapid escape of fumes from the jar indicates that it is filled. Collect some jars of the gas and close them with well-greased plates. In the wash-bottle is concentrated sulphuric acid to dry the

The reaction is very vigorous with the salt in fine crystals; to make the action slower compact masses obtained by fusing the salt are employed

In industry for the same mass of acid twice the mass of sodium chloride is employed and the mixture is heated much more strongly so that the sulphate of sodium without any hydrogen is produced.

Large quantities of hydrogen chloride are produced as a by-product in the preparation of sodium carbonate, in which common salt is first converted into sodium sulphate by the action of sulphuric acid.

Hydrogen chloride is a colourless gas heavier than air. It is extremely soluble in water, which dissolves almost 500 times its volume of the gas.

**Exp**—Open a jar of the gas under water to which some blue litmus solution has been added. Note that the water rushes up and fills the jar, while at the same time it turns red. 400 volumes of the gas dissolve in one volume of water.

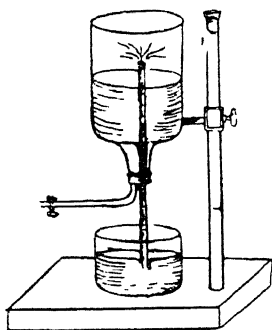


Fig. 52.

This experiment may be arranged in the form of a fountain (Fig. 52.) The bottle is filled with the gas and the tumbler contains water coloured blue with litmus solution. The water rises in the form of a fountain, becoming red. The experiment may be arranged also in the other form with the funnel (Fig. 50).

The gas fumes strongly in the air, as it combines with the aqueous vapour in the air.

The ordinary commercial hydrochloric acid has 421 grams of the real acid in 1180 grams. It is generally coloured yellow from the small quantities of iron derived from the apparatus in which it is prepared. It fumes in the air, because it slowly gives off the gas dissolved in it.

Gaseous hydrogen chloride gives in contact with ammonia gas abundant white fumes of ammonium chloride (*sal-ammoniac*).

**Exp.**—Put a few drops of liquor ammonia in a jar; after shaking it well close it with a glass plate, invert it and place it over a jar of hydrogen chloride. Now remove the plates between the jars, and keeping the two together invert the whole. Observe the thick white smoke formed within, condensing on the sides.

Hydrochloric acid acts in the cold on a large number of metals with evolution of hydrogen and the formation of metallic chlorides. This property is utilised in the preparation of hydrogen by the action of the acid upon zinc or iron.

Zinc + hydrogen chloride = zinc chloride + hydrogen.

The gas attacks most metals at a more or less high temperature, always forming chlorides with evolution of hydrogen.

We thus see that the hydrogen of hydrochloric acid is replaceable by metals.

Hydrochloric acid dissolves a large number of the oxides, sulphides and carbonates of metals always forming the chlorides of the corresponding metals.

Copper oxide + hydrochloric acid = zinc chloride +  
water.

Iron sulphide + hydrochloric acid = iron chloride +  
hydrogen sulphide.

Calcium carbonate + hydrochloric acid = calcium  
chloride + carbon dioxide + water.

All these experiments prove the presence of hydrogen in hydrogen chloride. Hence **hydrochloric acid gas is a compound of hydrogen**. What are its other components?

You may try to see what remains when hydrogen is removed from the gas by means of oxygen.

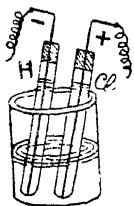
**Exp.**—Conduct a current of oxygen through hydrochloric acid and after drying the mixture lead it through a combustion tube filled with broken brick and heated very strongly; a greenish-yellow gas with a highly irritating smell and causing cough is produced.

The hydrogen of hydrochloric acid is oxidised on heating by compounds rich in oxygen (such as

manganese dioxide, potassium permanganate, lead peroxide) with evolution of chlorine.

**Exp.**—Heat some fragments of manganese dioxide or a few crystals of potassium permanganate in a test-tube with concentrated hydrochloric acid. Observe the greenish-yellow gas with irritating odour. Place a strip of blue litmus paper on the mouth of the tube. Note how the colour is destroyed and the paper becomes white. This gas is chlorine.

Manganese dioxide + hydrochloric acid = Manganese chloride + water + chlorine.



In the electrolysis of a concentrated solution of hydrochloric acid (Fig. 53) chlorine is liberated at the positive electrode.

Therefore the gaseous compound produced by the action of sulphuric acid upon sodium chloride is a compound of hydrogen and chlorine.

**Exp.**—Pass dry hydrochloric acid gas over sodium heated in a bulb tube. The sodium glows and becomes covered with a white crust, while at the end of the tube the presence of the hydrogen liberated can be proved by its burning. After the action has taken place sufficiently, allow the white crust to cool and dissolve it in water; evaporate the solution to crystallise. The crystals are small cubes and by testing them you recognise them to be common salt. In the hydrochloric acid gas the hydrogen is replaced by the metal sodium. Therefore common salt is a compound of sodium and chlorine.

The salts of hydrochloric acid or hydrogen chloride are called **chlorides**. The chlorides of most metals are soluble in water, and their solutions as well as hydrochloric acid give on the addition of a solution of silver

nitrate a white curdy precipitate of silver chloride which turns first purple, then black on exposure to light. Silver nitrate solution is a reagent for chlorides in solution.

---

## CHLORINE.

**Chlorine.**—For experimental purposes chlorine is most conveniently prepared as follows.

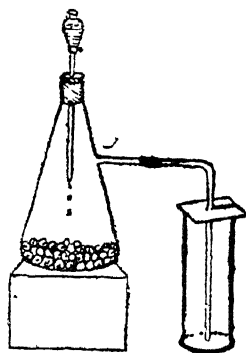


Fig. 54.

**Exp.** Fit up apparatus as shown in Fig. 54. Introduce into the flask enough potassium permanganate in small lumps so as to form a layer 1 c.m. deep. Commercial hydrochloric acid is allowed to drop down the dropping funnel. The gas is evolved without heating. Collect the gas-evolved in jars by downward displacement of air, and cover them with glass plates.

Potassium permanganate, a violet compound of potassium, manganese and oxygen is very rich in oxygen, and is a very powerful oxidising agent. The oxygen of this compound oxidises the hydrogen of the hydrochloric acid into water and liberates the chlorine. A dilute aqueous solution of potassium permanganate is often used as a mouth wash.

Instead of potassium permanganate manganese dioxide and concentrated hydrochloric acid, or a mixture of common salt, manganese dioxide and sulphuric acid, is employed. Then heat is required. In the second case the hydrogen chloride liberated by the action of sulphuric acid on common salt acts upon the manganese oxide to produce chlorine.

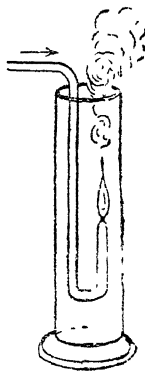
**N.B.** When chlorine is not collected, insert the exit tube in a beaker containing slaked lime. Do not allow the gas to escape into the room.

Observe that the gas is greenish and has an irritating odour. It irritates the throat and produces cough. It is heavier than air, the mass of 1 litre being 3.2 grams.

Chlorine is soluble in water and the solution is named **chlorine water**.

The most important property of chlorine is the ease with which it combines with hydrogen and the metals. Equal volumes of hydrogen and chlorine slowly combine to form hydrogen chloride under diffuse light. Direct sunlight makes the two gases combine swiftly with a violent explosion.

**Exp.**—Introduce a jet of burning hydrogen into a jar of chlorine (Fig. 52). The flame continues to burn with a pale greenish flame; white fumes escape from the jar owing to the formation of hydrogen chloride. This represents the synthesis of hydrogen chloride, proving that it is a compound of hydrogen and chlorine.



Also many metals burn in chlorine and are transformed into chlorides.

**Exp.**—Throw into a jar of chlorine some Dutch metal (copper leaf). The metal immediately takes fire. In the jar remains a greenish-white solid, which, on shaking with water, gives a green solution of copper chloride.

Copper + chlorine = copper chloride.

Fig. 55. 2. Similarly drop finely powdered antimony into another jar. Each particle takes fire as it falls. Antimony + chlorine = antimony chloride.

Melt a piece of sodium on a deflagrating spoon and lower it into another jar. The sodium combines with the chlorine forming white sodium chloride and finally bursts into flame.

Sodium + chlorine = sodium chloride.



Chlorine attacks a large number of compounds of hydrogen.

It slowly decomposes water under the influence of light liberating oxygen.

Chlorine + water = hydrogen chloride + oxygen.

This is why chlorine water has to be preserved in black bottles.

If water contains in solution substances easily oxidisable, they are at once oxidised by chlorine.

Chlorine passed into a solution of sulphurous acid at once oxidises it into sulphuric acid

All organic colouring matters are readily decolorised by chlorine in the presence of water.

**Exp.**—Throw into a jar of dry chlorine two strips of indigo cloth, one wet and the other dry, a moist rose, a moist green leaf, a piece of moistened printer paper. All are bleached, except the dry indigo cloth and the printed paper. **Chlorine is a powerful bleaching agent.**

Indigo, the flower, the leaf are all decolorised by chlorine in the presence of water, that is to say, are decolorised by the oxygen liberated by the combination of chlorine with the hydrogen of the water, which oxidises and destroys the organic colouring matters. The printed paper is unaffected, because printers' ink contains lamp-black, which is an inorganic pigment.

All organic substances contain hydrogen and are more or less violently attacked by chlorine.

**Exp.**—Dip a piece of filter paper in warm oil of turpentine and quickly drop it into a jar of chlorine. It at once catches fire and burns depositing a large amount of soot.

**Exp**—Attach a lighted taper to a deflagrating spoon and introduce it into a jar of chlorine. The taper continues to burn with a diminished red sooty flame. The wax of the taper and the oil of turpentine are compounds of carbon and hydrogen. The chlorine

combines with the hydrogen in them to form hydrogen chloride, liberating the carbon as soot. The presence of hydrogen chloride can be proved with blue litmus paper or ammonia.

As cork and rubber are organic substances, chlorine attacks them, rubber more than cork.

**Bleaching powder**, called often **chloride of lime**, is obtained by subjecting freshly slaked lime to the action of chlorine till it will not take up any more. On treatment with dilute acids bleaching powder yields chlorine. Cotton goods are bleached by first dipping them in a clear solution of bleaching powder, then in very dilute sulphuric or hydrochloric acid and finally washed. The solution of bleaching powder should not be used in bleaching silk or woolen goods, because chlorine attacks their fibres and weakens them.

Bleaching powder is used for chlorinating water to destroy organic impurities and cholera germs.

---

## XXVI

### IODINE.

Chlorine liberates **iodine** from solutions of iodides.

**Exp.**—Examine some **potassium iodide**, a salt used in medicine. It consists of cubical crystals resembling common salt in appearance. It dissolves in water. Dissolve some crystals in a test tube and add to the solution drop by drop some chlorine water. The liquid becomes dark brown first, as the liberated iodine dissolves in excess of potassium iodide. On adding more chlorine water all the iodide is decomposed, and the iodine set free swims as a dark-grey solid in the liquid, as iodine is very slightly soluble in water.

1. Take a portion of this liquid in another test tube and add some alcohol. On shaking the test tube a brown liquid floats on the surface of a clear liquid. Alcohol dissolves the dark grey solid and being lighter than water floats on its surface.

2. To a second portion of the brown liquid add some drops of carbon disulphide. On shaking the carbon disulphide takes up the brown substance and collects at the bottom of the tube as a violet-coloured liquid, since carbon disulphide is heavier than water.

3. Put the remainder of the liquid in a round-bottomed flask. Evaporate the liquid by gentle heat; then heat somewhat strongly. The brown stuff at once changes into a dense violet vapour. On allowing the flask to cool fine dark violet crystals with shining facets are deposited on the sides of the flask.

Iodine is sold in the form of small crystalline plates with a dark violet colour and feeble metallic lustre. It does not perceptibly dissolve in water, but easily dissolves in water in the presence of potassium

iodide. It dissolves easily in carbon disulphide to a violet solution and in alcohol to a brown solution. An alcoholic solution of iodine is employed in medicine as **tincture of iodine**.

**Exp.**—Evaporate some tincture of iodine in a round-bottomed flask and note what happens.

A striking property of free iodine is the blue colour it imparts to starch, which occurs as fine white granules in the grains of cereals (rice, wheat) and in tubers (potato).

**Exp.**—Make a thin paste by boiling a little starch with water. Add to the cold liquid a little alcoholic solution of iodine. Note the dark-blue colour produced, due to the formation of a compound of starch with iodine. Starch is a reagent for iodine.

If starch paste is a reagent for free iodine, a solution of iodine is equally a reagent for starch.

**Exp.**—Cut some sections of a potato. Put on each of them a drop of the iodine solution. Note the blue colour, which indicates the presence of starch in the potato.

**Exp.**—Add to starch paste a solution of potassium iodide. There is no change of colour, as there is no free iodine. Now add a trace of chlorine water; at once a blue colour appears, owing to the liberation of free iodine from the iodide by chlorine.

Iodine is prepared industrially from **kelp**, the ash of seaweeds, in which it occurs in the form of sodium iodide. The ash is first dissolved in water and partially evaporated, when the difficultly soluble sodium chloride crystallises out. To the mother liquor left behind, which contains sodium iodide, chlorine water is added, which sets the iodine free. From this liquid iodine is separated by evaporation and sublimation.

---

## XXVII.

### NITRE

**Nitre**, or **saltpetre**, is a white crystalline salt with a bitter taste. It is very soluble in water and from its solution it crystallizes out in long needles or prisms.

**Exp.**—Introduce a little powdered nitre on an asbestos thread into a non-luminous flame. The flame becomes violet-coloured. Nitre therefore contains potassium.

**Exp.**—Heat some nitre in a test tube till the salt melts. Drop into the molten mass a piece of charcoal and heat strongly. Suddenly the charcoal begins to glow strongly, as it does in oxygen. Nitre gives off oxygen at higher temperatures and must therefore contain oxygen.

**Exp.**—Pour some drops of concentrated sulphuric acid on a crystal of nitre in a test tube. Note the evolution of brown fumes of nitrogen peroxide. Nitrogen peroxide is a compound of nitrogen, and sulphuric acid contains no nitrogen. Hence nitre contains nitrogen.

Nitre is therefore a compound of potassium, oxygen and nitrogen. Its chemical name is potassium nitrate.

Nitre occurs in all places where animal matters, which are always rich in nitrogen, undergo putrefaction in the presence of potash and slaked lime. First ammonia is produced, but when this comes in contact with the air and potash, or lime, microbes oxidise it into nitric acid, and this produces the nitrate of potassium or calcium and water. If calcium nitrate is formed, it is decomposed with potash, whereby potassium nitrate and calcium carbonate are formed. The

latter being insoluble in water, potassium nitrate is separated by solution, evaporation and crystallization.

At the present day large quantities of potassium nitrate are obtained by mixing hot concentrated solutions of sodium nitrate (Chile saltpetre) and potassium chloride (abraum salts), both of which occur in nature in abundance. Sodium chloride, being less soluble than potassium nitrate, separates on evaporating the hot mixture, and from the liquid left behind potassium nitrate crystallizes out on cooling.

Sodium nitrate + potassium chloride = sodium chloride + potassium nitrate

Sodium nitrate occurs in large beds in Chile in South America.

Large quantities of nitre are used in the manufacture of gun-powder (black powder). Ordinary gun powder consists of sulphur, charcoal (both easily inflammable materials) and nitre. Nitre provides the oxygen necessary for the complete combustion of the other two substances in a limited space. The large amount of gas produced by the combustion, being confined in a small volume, exerts an enormous pressure, which is still further increased by the heat generated in the combustion, and this pressure drives the projectile out of the gun with great force.

Dry sodium nitrate possesses the same oxidizing properties as nitre, but being deliquescent does not find employment in the manufacture of gun-powder.

**Exp.**—Dip a strip of filter paper in a solution of nitre dry it and burn it. Note what takes place and explain it. This is how **touch-paper** is made.

---

## NITRIC ACID. NITRATES.

**Exp.**—Concentrated nitric acid is easily prepared by the action of concentrated sulphuric acid on potassium

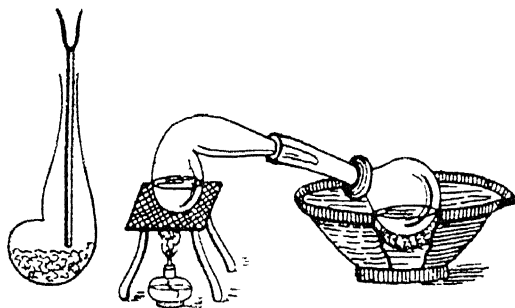


Fig. 56

nitrate or sodium nitrate with the apparatus shown in Fig. 56. In a retort place some grams of nitre and cover it with an equal mass of concentrated sulphuric

acid by pouring the acid through a long thistle funnel so that the sides of the retort are not touched by the acid. The neck of the retort is fitted into a glass receiver, which is placed in a basin of cold water. To keep the receiver under water its neck is loaded with a ring of lead. At first the receiver is filled with reddish fumes, as sulphuric acid removes water from the nitric acid formed at first. Then these fumes disappear and a clear liquid collects in the receiver. Fumes reappear in the retort at the end of the reaction, because the temperature then rises.

In industry nitric acid is manufactured by the same process, but the nitrate of potassium is replaced by the native nitrate of sodium (Chile salt-petre) as the mass of nitric acid produced with a given mass of sulphuric acid needs a smaller mass of sodium nitrate than of potassium nitrate.

Commercial nitric acid contains the real acid hydrogen nitrate, together with varying quantities of water. The pure acid is colourless and has a specific gravity of 1.5. The ordinary acid contains about 46% of its mass of water; it is coloured pale yellow on account of traces of nitrogen peroxide. It fumes in the air, because it emits vapours which combine with the moisture present in the air.

Pure nitric acid is decomposed when heated into peroxide of nitrogen, water and oxygen.

Nitric acid = nitrogen peroxide + water + oxygen.

This tendency to give up oxygen makes nitric acid an energetic oxidising agent. It is very dangerous on account of its corrosive action on the skin. It attacks violently all substances capable of combining with oxygen.

**Exp.**—Heat a porcelain dish with some charcoal till the charcoal is nearly red-hot. Place the dish for safety in a large beaker half filled with sand. Carefully pour over the charcoal some concentrated nitric acid. The charcoal is inflamed owing to the oxidising action of the acid.

Metals are transformed into nitrates by nitric acid, the acid at the same time being reduced to compounds of nitrogen with less oxygen. Thus with metallic copper we get brown nitrous oxide and a greenish blue liquid (copper nitrate).

Gold and platinum are the only metals unaffected by nitric acid but are dissolved in a mixture of 3 volumes of concentrated hydrochloric acid and 1 volume of concentrated nitric acid. This mixture has been called *aqua regia*, because it dissolves gold, the king of metals, while nitric acid has been called *aqua fortis*.

In no case by the action of nitric acid on metals do we obtain hydrogen, as the hydrogen evolved is completely oxidised by nitric acid.



Nitric acid oxidises a large number of compounds. You have already seen how it changes sulphurous anhydride into sulphuric anhydride.

Nitric acid destroys organic substances. Indigo is decolorised. Skins, feathers, wool, straw are all affected. A drop falling on a cloth produces an indelible yellow stain.

Nitric acid is specially used for the preparation of explosives, such as nitro-glycerine, gun-cotton, trinitro-toluene. Gun-cotton is the material from which **artificial silk** is produced. A compound similar to gun-cotton treated with alcohol and camphor produces **celluloid**, largely employed for combs, handles, small cases, photographic films, &c.

When nitric acid dissolves metals, their oxides, or their carbonates, nitrates are produced. All metallic nitrates are soluble in water.

Besides nitre and sodium nitrate, the following nitrates are important.

**Ammonium nitrate**, produced by the action of nitric acid on ammonia, is used in smokeless and flameless explosives. Ammonium nitrate when heated breaks up into nitrous oxide (employed as a mild anaesthetic) and water.

**Barium and strontium nitrates** are employed in fireworks for "green fire" and "red fire" respectively.

**Silver nitrate** is used in photography and in making indelible inks.

All nitrates are decomposed by concentrated sulphuric acid with liberation of nitric acid, just in the same manner as metallic chlorides are decomposed by the same acid with evolution of hydrogen chloride.

---

## BONE ASH. PHOSPHORUS.

Bones contain nearly thirty three percent by mass of animal matter ; the remainder is a mixture of calcium phosphate(four-fifths) and calcium carbonate(one-fifth). When a dry bone is strongly heated with a good supply of air, all the organic matter is burnt away, and the residue, though it may retain the form of the bone, is white and very fragile. This ash is called **bone ash**.

**Exp.**—Heat a piece of dry bone very strongly (say, in a charcoal furnace). Observe the white calcined bone left behind. The ash consists of a mixture of calcium and magnesium phosphates and calcium carbonate. Treat the ash with dilute hydrochloric acid and test the gas for carbon dioxide.

If bone ash is treated with dilute sulphuric acid, **phosphoric acid** goes into solution, while insoluble sulphate of calcium remains behind.

If moderately concentrated sulphuric acid is employed, the temperature rises during the reaction and leaves a crystalline mass mixed up of calcium sulphate, acid phosphate of calcium, and a little free phosphoric acid. This mixture is called **super-phosphate** and is used as manure.

The solution of phosphoric acid is syrupy, and when sufficiently concentrated crystallises. When strongly heated, it loses water and is converted into glacial phosphoric acid. When this is heated with charcoal (carbon) to a high temperature, the element phosphorus is produced.

Phosphoric acid (syrupy) = glacial phosphoric acid  
+ water

Glacial phosphoric acid + carbon = phosphorus +  
water + carbon monoxide.

A mixture of phosphoric acid concentrated to the consistency of a syrup and powdered charcoal is dried and then heated to  $900^{\circ}\text{C}$  in earthenware retorts. At this temperature the carbon of the charcoal takes up the oxygen of the acid to form carbon monoxide and the phosphorus set free is changed into vapour, which escapes from the retorts with the carbon monoxide through chests full of water maintained at a temperature higher than  $50^{\circ}\text{C}$ . In these chests the vapours of phosphorus liquefy, while carbon monoxide escapes. The molten phosphorus is purified under hot water and moulded into sticks under tepid water which is allowed to cool slowly.

Calcium phosphate occurs in nature as the mineral apatite, which is used both as a fertiliser and as a source of phosphorus.

**Phosphorus** is a whitish yellow solid at ordinary temperatures with a peculiar smell. Its specific gravity is 1.8. It melts at  $44^{\circ}\text{C}$ . It is insoluble in water, but soluble in carbon disulphide. The solution of phosphorus in carbon disulphide is highly inflammable.

**Exp.**—Dissolve a small bit of phosphorus in some carbon disulphide and pour the solution on filter paper. The solvent rapidly evaporates leaving the phosphorus behind in a finely divided state, and this at once takes fire spontaneously.

Yellow phosphorus is highly inflammable. On account of the readiness with which it catches fire it has always to be kept and handled under water. It burns in air as soon as it melts, giving off white fumes of **phosphoric anhydride** with a garlic odour. It burns in oxygen even under water.

**Exp.**—Place some phosphorus in a gas jar and cover it with water hot enough to melt it. Then by means of a tube reaching to the bottom of the jar lead oxygen close to the molten phosphorus. Each bubble of oxygen, as it comes in contact with the phosphorus, gives rise to a flash.

Yellow phosphorus exposed to air slowly oxidises in the cold, emitting white fumes. The luminosity of phosphorus in the dark is due to its slow oxidation. Phosphorus (light-bearer) owes its name to this characteristic. The luminous trace which a lucifer match leaves on rubbing it on a hard body is due to small particles of phosphorus left behind, which oxidise in the air.

Phosphoric anhydride dissolves in water with a hissing sound, showing its great avidity for water. It readily absorbs moisture from the air and becomes wet. It finds employment as a drying agent in the laboratory. The aqueous solution of phosphoric anhydride contains phosphoric acid, whose salts are named **phosphates**.

On account of its readiness to combine with oxygen, phosphorus forms a powerful reducing agent. It reduces nitric acid forming phosphoric acid (dangerous reaction). Finely divided phosphorus combines with the oxygen of potassium chlorate with explosive violence. This may be seen by pouring some solution of phosphorus in carbon disulphide on some heaped up potassium chlorate powder.

Phosphorus and its vapour are highly poisonous.

When heated for a long time in the absence of air, yellow phosphorus is transformed into a dark red powder, called **red phosphorus**. This is denser than yellow phosphorus. It does not dissolve in carbon disulphide. It does not melt. It does not glow in the dark. It does not burn in air, till it is heated to  $260^{\circ}\text{C}$ . It consists of crystalline particles, though called **amorphous**. Red phosphorus is not poisonous.

Its chemical properties are, however, the same as those of ordinary phosphorus, but its reactions are less vigorous. 62 grams of phosphorus, yellow or red, combine with 80 grams of oxygen, yielding 142 grams of phosphoric anhydride. Therefore red phosphorus is regarded only as a physical variety of the same element as yellow phosphorus.

**Friction matches**, formerly in use, consist of wooden strips coated with sulphur or paraffin and tipped with a paste containing yellow phosphorus and potassium chlorate. As yellow phosphorus produces a serious disease in workers handling it, friction matches have been displaced by **safety matches**. The tips of these contain no phosphorus, but a paste of antimony sulphide, potassium chlorate and red lead; these take fire only on rubbing over an igniting surface coated with a paste of powdered glass and red phosphorus placed on the sides of the box. Friction matches are still made with a non-poisonous sulphide of phosphorus.

**Exp.**—Take a large crystal of potassium chlorate and with pincers rub it on the igniting surface of a match box. Note the flash produced.

---

## XXX

### ACIDS. SALTS. BASES

**Acids.**—Of the many chemical compounds you have come across in your study, some, like carbonic acid, sulphurous acid, sulphuric acid, hydrochloric acid, nitric acid, phosphoric acid, bear in common the name **acid**, because they have a sour taste and turn blue litmus red. This acid character was supposed by Lavoisier to be due to the presence of oxygen in them, but it is not so, as we know acids, like hydro-chloric acid (hydrogen chloride), hydro-sulphuric acid (hydrogen sulphide), which do not contain oxygen. Again, there is an acid which is insoluble in water and has no sour taste. Further, sulphur dioxide, carbon dioxide, hydrogen sulphide, taste sour and turn blue litmus red only in aqueous solution ; in perfectly dry condition they are not acid. Which element then is an essential constituent of acids ?

We have seen that the element common to hydrochloric acid and hydro-sulphuric acid is hydrogen. By the action of water on iron or zinc at ordinary temperature no hydrogen is liberated. On the addition of sulphuric acid hydrogen is liberated. This hydrogen does not come from the metals, as they are elements. The evolution of hydrogen begins just when sulphuric acid is added ; it increases when the quantity of acid is increased ; it ceases when the acid is used up. Hence sulphuric acid also contains hydrogen. It is more difficult to prove the presence of hydrogen in nitric acid. If, however, the vapour of nitric acid is passed through a white hot porcelain tube, it is decomposed into water, oxygen and nitrogen. Water indicates the presence of hydrogen as a component of nitric acid. We therefore conclude that :

**All acids are compounds of hydrogen.**

**Salts.**—In a boiling flask put some zinc and hydrochloric acid; in another put some iron and dilute sulphuric acid. Hydrogen is evolved in both cases. Gradually the metals disappear, the evolution of gas becomes feebler, and ceases when all trace of the metals has disappeared. What has become of the metals?

In order to find this filter the liquids; the first filtrate is colourless, the second green. Evaporate the liquids in flat dishes. In the first small colourless crystals make their appearance, in the second appear large green crystals. As both of these stuffs resemble common salt in being soluble in water and crystallising out of their aqueous solutions, they are named **salts**. The first is produced by zinc taking the place of hydrogen in hydrochloric acid and the second by iron taking the place of hydrogen in sulphuric acid. Therefore :

**Acids are compounds of hydrogen, in which the hydrogen can be replaced by a metal.**

**A salt is produced when the hydrogen of an acid is replaced by a metal.**

A special feature of salts is that their aqueous solutions conduct a current of electricity and are spoken of as **electrolytes**. In general salts have a neutral reaction, but some salts, like blue vitriol, alum, have an acid reaction, and some, like potassium carbonate and borax (sodium borate), have an alkaline reaction.

**Bases**—By dissolving zinc in hydrochloric acid we obtain a zinc salt, which contains besides the metal zinc the residue (radicle, ion) of the acid, namely, chloride; it has therefore the name zinc chloride. Similarly by dissolving iron in dilute sulphuric acid we obtain a salt which contains besides the metal iron the residue (radicle) of sulphuric acid and is called iron sulphate. Of what does this acid residue (radicle) consist? When we pour concentrated sulphuric acid on iron in a boiling flask, we perceive first no action of the acid on the metal. But on heating the mixture sulphur dioxide is evolved, which is at once recognised by its smell. This

shows that sulphuric acid contains sulphur and oxygen. Therefore, when the hydrogen is replaced by the metal iron, the salt formed must contain besides the metal also sulphur and oxygen.

We can also obtain salts in another way. If hydrochloric acid or dilute sulphuric acid is poured on zinc oxide or iron oxide, no gas is evolved but the oxide dissolves in the acid. On evaporating the solutions we obtain in the case of zinc oxide and sulphuric acid colourless crystals and in the case of iron oxide and sulphuric acid green crystals. Thus we obtain the same salts as we do when the metals zinc and iron are dissolved in dilute sulphuric acid. But no gas is evolved, as the hydrogen of the acid combines with the oxygen of the oxide to form water. Metal oxides form the basis for the formation of salts and have therefore been called **bases**.

**Bases combine with acids to form salts, producing at the same time water.**

**Alkalies.**—If bases are soluble in water, they turn red litmus blue; they have the taste of lye and have a corroding action on organic substances. Such bases are called **alkalies**. Alkalies turn turmeric red; hence the action of slaked lime on turmeric.

**Exp.**—Burn in a spoon a piece of metallic sodium. The white residue is sodium oxide, which easily dissolves in water and exhibits the above-mentioned characteristics. Sodium oxide is an alkali, a strong base. Similarly from potassium we obtain potassium oxide, which is likewise an alkali.

Place a strip of red litmus paper in a solution of oxide of sodium contained in a dish. The paper at once turns blue. Now add from a burette drop by drop hydrochloric acid to the dish, till the blue litmus paper just turns purple, which indicates that there is neither free alkali nor free acid. The solution is said to be **neutral**, and has no action on either blue litmus or red



litmus. The acid has **neutralised** the alkali, and produced a chemical compound.

Evaporate the solution. We obtain small cubical crystals of a salt, which we recognise by its taste to be common salt, sodium chloride.

How can we obtain potassium chloride, sodium sulphate, potassium nitrate?

In the naming of salts the aim is to indicate the metal first and then the radicle of the acid from which it is formed. Thus the salts of :

Hydrochloric acid	are named	chlorides,	<i>e.g.</i>	sodium chloride
Sulphuric	„	„	„	„ sulphates „ copper sulphate
Carbonic	„	„	„	„ carbonates „ calcium carbonate.

**Exp.**—Introduce a bright knife-blade into a solution of copper sulphate (blue vitriol). In a short time the blade is coated with a red deposit, which is at once recognised to be copper. Now remove the knife-blade and add some iron filings and warm gently, the blue colour becomes fainter. Add more iron filings if necessary, till the blue colour disappears entirely. Filter the nearly colourless liquid into an evaporating dish, and allow it to crystallise by slow evaporation. The green crystals are recognised to be identical with those of green vitriol (ferrous sulphate). So iron has entered into chemical combination displacing copper.

Copper sulphate + iron = iron sulphate + copper.

**Exp.**—Suspend a strip of sheet zinc in a solution of lead acetate (a highly poisonous substance which on account of its sweet taste is called **sugar of lead**) contained in a beaker. After some days shining crystals of metallic lead collect around the sheet. The zinc acts slowly on the lead compound in the solution, displaces metallic lead, and forms zinc acetate.

Lead acetate + zinc = zinc acetate + lead.

These experiments prove that some metals are able to displace other metals from their salts.

You at once recognise a similarity between these and the reactions in which metals displace hydrogen from acids.

Iron + sulphuric acid = iron sulphate + hydrogen

Zinc + hydrochloric acid = zinc chloride + hydrogen

Hence we may regard acids as salts of hydrogen. Thus sulphuric acid is called hydrogen sulphate; hydrochloric acid hydrogen chloride; nitric acid hydrogen nitrate.

The termination **ate** in the name of a salt indicates the presence of oxygen as a component and the termination **ide** (except in the case of oxides) implies the absence of oxygen. Thus potassium chloride contains only potassium and chlorine, while potassium chlorate contains potassium, chlorine and oxygen.

---

## XXXI.

### MERCURY. LEAD. TIN.

**Mercury, or quicksilver** (Latin, *hydrargyrum*) occurs in nature in the free state to some extent. More frequently it occurs in combination with sulphur as the dark-red sulphide **cinnabar**.

When cinnabar is heated in a current of air, its component sulphur is easily oxidised to sulphur dioxide. If it is heated with iron, the sulphur combines with the iron to form iron sulphide. The escaping vapour of mercury is condensed in receivers, and after pressing the mercury through fine leather it is again distilled.

Mercury is the only metal, which is liquid at ordinary temperatures. It has the appearance of molten silver. Its specific gravity is 13.59. All metals, with the exception of gold and platinum, and all minerals float on mercury. It solidifies at  $-39^{\circ}\text{C}$ . and boils at  $359^{\circ}\text{C}$ . It gives off vapour at all temperatures, which is poisonous.

When heated continuously in contact with air at a temperature close to its boiling point, it becomes changed into the **red oxide**. Mercury easily dissolves in nitric acid forming **mercurous nitrate** if the metal is in excess, or, **mercuric nitrate** if the acid is in excess. Its chlorides, **corrosive sublimate** (mercuric) and **calomel** (mercurous), are important, as both of them are used in medicine. Its sulphide, **vermillion**, is a fine scarlet pigment. The white rods, which burn producing a long snake-like ash, and are sold under the name of **Pharaoh's serpents**, contain a compound of mercury. All compounds of mercury are poisonous.

Mercury and mercury compounds in acid solution produce on copper a deposit, which on rubbing becomes

bright like silver. Copper displaces mercury from its compounds and the mercury liberated forms an amalgam with copper. By this test the presence of mercury may be detected.

**Lead.**—The most important ore of lead is **galena**, a sulphide with a bluish-grey colour and metallic lustre and often crystalline.

Lead (Latin, *plumbum*) is obtained from galena by first roasting it and then melting the residues obtained by roasting. On roasting galena becomes partly converted into lead oxide and lead sulphate by the action of the oxygen of the air. On melting the residue the oxygen present in the lead oxide and lead sulphate react upon the sulphur of the unchanged galena and change it into sulphur dioxide, which escapes, leaving metallic lead behind.

Lead sulphide + lead oxide = lead + sulphur dioxide.

Lead sulphide + lead sulphate = lead + sulphur dioxide

Lead is a metal of bluish-grey colour. Its fresh surface is bright, but like zinc it very soon loses its lustre in air, becoming coated with a layer of oxide. It is soft and malleable but not tenacious. Its specific gravity is 11.4. It melts at 330°C.

Lead forms useful alloys with some metals : type-metal (with tin and antimony); plumber's solder (with tin); shot-metal (with arsenic).

When lead is heated in air without melting, it forms a yellow amorphous powder (*massicot*); molten lead yields a bright yellow crystalline mass, which is finely powdered and sold under the name of **litharge**. Litharge, when heated to 350°C in air, absorbs oxygen and changes into **minium** or **red lead**, which on heating to 400°C decomposes into litharge and oxygen.

Lead dissolves easily in nitric acid, but not in hydrochloric and sulphuric acids, because the first formed lead chloride and lead sulphate, being insoluble, cover the metal and prevent further action. In the

presence of air organic acids, like acetic acid, dissolve lead forming poisonous compounds, Therefore lead is unsuitable for cooking vessels.

By the action of nitric acid red lead is still further oxidised to the snuff-coloured oxide, **lead peroxide**, which is often employed in the tips of safety matches.

The pigment, known as **white lead**, is a compound of lead carbonate and lead hydroxide, which is formed by the action of carbon dioxide on lead acetate. Lead acetate is highly poisonous and is known on account of its sweetish taste as **sugar of lead**.

**Tin** -Tin does not occur in the free state in nature, but occurs in combination with oxygen as the mineral **tin-stone** (cassiterite).

From this ore tin is extracted by roasting it to drive off any sulphur and arsenic present and then mixing it with coal and heating the mixture to redness.

Tin (*stannum*) is a white metal with a bright lustre and specific gravity 7.3. When a rod of tin is bent, it gives out a cracking sound. It melts at 235°C., and when melted, it becomes easily covered with a grey scum which is **stannic oxide**. Tin is very malleable and easily beaten into thin leaf. As tin does not tarnish in air, it is used to coat plates of iron and articles of copper.

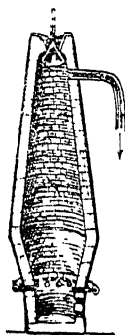
Hydrochloric acid dissolves tin with evolution of hydrogen forming stannous chloride. Concentrated sulphuric acid converts it on heating into a sulphate with evolution of sulphur dioxide. Dilute nitric acid dissolves tin with evolution of nitrogen oxide, forming a nitrate. Concentrated nitric acid on the contrary converts it into an insoluble oxide; concentrated solutions of caustic soda and caustic potash dissolve it on heating with evolution of hydrogen.

---

## IRON. COPPER. ZINC.

**Iron** (Latin, *ferrum*) occurs native only in meteorites. but is found in combination with other elements in very large quantities. The most important ores are the oxides, **red haematite** and **magnetite**, and the carbonate, **clay-ironstone**.

These ores are smelted by heating them with carbon (charcoal, coal, or coke) to a high temperature in a huge upright furnace, called a **blast furnace** (Fig 54), because a blast of air is forced through the melted mass to facilitate fusion and chemical change.



Mixed with iron ores is always found much earthy material containing silica and silicates. These being infusible, limestone is added to render them fusible. Into the blast furnace are dropped the fuel (charcoal, coal, or coke), the lime-stone, and the ore so as to form alternate layers. As the charge sinks fresh loads are dropped through the opening at the top. The fire once started is kept burning for months or years. Hot air is driven from below through blast pipes. Oxygen combines with the carbon of the fuel forming oxides of carbon. The carbon reduces the ore and the molten iron falls in drops to the bottom. The intense heat causes the combination of the earthy matter of the ore with the lime of the lime-stone to form a slag (an impure glass), which floats on the molten mass of iron collecting at the bottom. The molten iron is drawn off into sand moulds. This is cast iron (or pig iron). From cast iron are made steel and wrought iron. The slag hardens on cooling and is used to manufacture cement and fertilisers.

**Iron pyrites** is a sulphide of iron, found in abundance in the surface of the earth. On account of the large percentage of sulphur in it, it is not used for extracting iron.

Of the three varieties of iron in ordinary use wrought iron contains the smallest percentage of carbon and melts between  $1800^{\circ}$  and  $2250^{\circ}\text{C}$ . Steel is made either by adding carbon to wrought iron or by removing carbon from cast iron.

Steel is **hardened** by heating it to a high temperature and rapidly cooling it. By re-heating hardened steel to a temperature much lower than that used for hardening it and cooling it slowly, steel becomes **tempered**. Tempered steel is much more elastic and less brittle than hardened steel.

Dry air has no action upon iron at ordinary temperatures. Moist air, containing carbon dioxide, produces the rusting of iron. Iron rust is a complex compound. Rusting proceeds rapidly because the film of rust is not compact enough to protect the metal. When iron rust is heated strongly, it loses water and becomes converted into the dark red oxide (ferric oxide), which, under the name of **rouge**, is used for polishing gold and silver articles. Iron dissolves easily in hydrochloric acid and dilute sulphuric acid with evolution of hydrogen. Green vitriol (or copperas) is produced in the latter case. Green vitriol is obtained on a large scale by the action of air and moisture on iron pyrites. Green vitriol (ferrous sulphate) is used in making ink and some pigments. It is also used as a mordant for fixing colours on cloth. **Ferric chloride** is used in medicine.

**Copper** (Latin, *cuprum*) occurs native in extensive mines in America. The chief ores used for extraction are **red copper ore**, a red oxide, and **malachite**, a green carbonate.

Out of the red copper ore and malachite copper is extracted by reduction with carbon. Pure copper is obtained by electrolytic processes.

Copper has a red colour, which is clearly seen when a copper coin is cleaned with tamarind or dilute nitric acid. It is a good conductor of heat and electricity. It is malleable and ductile.

When bright copper is slightly heated in air, it oxidises slightly and becomes covered with a brownish red layer of oxide, which, on stronger heating, takes up more oxygen and becomes converted into the black oxide. In moist air copper becomes coated with a layer of green **verdigris** or carbonate.

Copper dissolves easily in nitric acid, but only on heating in concentrated sulphuric and hydrochloric acids. Pure copper sulphate is obtained by dissolving copper or copper oxide in sulphuric acid.

Solutions of copper salts give a deep blue solution on the addition of a solution of ammonia. Red copper is deposited on clean iron dipped into them.

Copper and copper compounds, moistened with concentrated hydrochloric acid and introduced on asbestos thread into a non-luminous flame, impart a green colour to the flame.

All copper compounds are poisonous. Hence all copper vessels used for cooking purposes must be well tinned.

How can the presence of copper in a conserve be detected? What is blue vitriol?

**Zinc** (Latin, *zincum*) occurs in nature as carbonate in **calamine** and as sulphide in **zinc blende**.

By roasting, that is, heating strongly with a good supply of air, both these ores are converted into zinc oxide. The zinc oxide thus obtained is mixed with carbon and heated in retorts when the oxide becomes reduced to metallic zinc, which distils over.

Zinc is a bluish-white metal, which can take a good polish, but loses its brightness in air becoming covered with a grey coating of oxide. It melts at  $420^{\circ}\text{C}$ . At ordinary temperatures it is brittle and coarsely crystalline in fracture. When heated a little above  $100^{\circ}\text{C}$  it becomes so soft that it can be hammered



and rolled. At about  $300^{\circ}\text{C}$ . zinc becomes so brittle that it can be ground to powder. By pouring a stream of molten zinc in water we get **granulated zinc**, that is, zinc in the form of small grains.

Zinc forms important alloys : brass (with copper) ; german silver (with copper and nickel) ; bronze (with copper and tin). To protect iron from rusting iron articles are coated with zinc by plunging them in molten zinc. The iron is then said to be **galvanised**.

The vapour of the metal burns with a bluish-green flame, producing a white smoke, which consists of zinc oxide, used as a paint under the name of zinc white.

Zinc dissolves in dilute sulphuric acid and hydrochloric acid with evolution of hydrogen forming respectively zinc sulphate and zinc chloride. Zinc chloride is very deliquescent. Crystals of zinc sulphate contain water of crystallisation.

---

## XXXIII

### ALUMINIUM.

**Aluminium** occurs in nature in the combined state. It occurs as silicate in all varieties of **clay** ; and as fluoride in **cryolite** ; as oxide in **corundum**, **ruby**, **sapphire**.

It is prepared on a large scale by the electrolysis of aluminium oxide which is dissolved in molten cryolite.

Aluminium is a silver-white, malleable metal of specific gravity 2.6. It melts at 660° without oxidation. It vaporises at 1800°C. and the vapour burns in the air forming aluminium oxide.

Owing to its lightness and resistance to the action of the air it is used in the manufacture of many articles of ordinary use. Aluminium leaf serves as a substitute for silver leaf, especially as it is not blackened by hydrogen sulphide. Aluminium powder is used in making metal varnishes. Mixed with ammonium nitrate the powder is used as an explosive. Aluminium forms many useful alloys with copper (aluminium bronze) and magnesium (magnalium).

Aluminium in powder or thin leaf burns easily in air with a dazzling light and decomposes boiling water with evolution of hydrogen. At its melting point metallic aluminium is an energetic reducing agent. Aluminium powder reduces the oxides of iron, chromium, manganese. It dissolves in hydrochloric acid with evolution of hydrogen. Dilute sulphuric and nitric acids attack it only very slowly ; in hot concentrated sulphuric acid it dissolves with evolution of sulphur dioxide. In alkalis it dissolves easily with evolution of hydrogen. Many salts, including common salt, act on aluminium.

The best known salt of aluminium is **alum**, a white soluble salt which crystallises out of its aqueous solution in octahedra (Fig. 58).

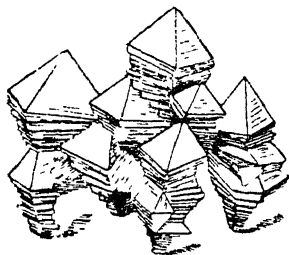


Fig. 58.

It consists of aluminium sulphate combined with potassium sulphate. It is employed in medicine, dyeing, tanning and paper manufacture. On being heated the crystals of alum give up their water of crystallisation, leaving a white porous mass, known as burnt alum. Alum is astringent and its solution is used

as a gargle.

**Clay** is a hydrated silicate of aluminium. **Pottery** is made of clay. Clay being plastic is moulded to desired shapes, which on strong heating turn into hard non-plastic **earthenware**. The purest clay is white **kaolin**, which is used for making **porcelain** or **chinaware**. Pottery is **glazed** to make it non-porous. **Bricks** and **tiles** are made from a common clay consisting of sand and clay, or marl. **Marl** is a mixture of clay, sand, and a loose form of calcium carbonate (limestone or shells) together with small quantities of iron oxide, lime and soda. The calcium carbonate makes the bricks hard and non-porous when baked. The red colour of ordinary bricks is due to the iron oxide. **Fire bricks** are made from **fire clay**, which is composed of kaolin mixed with sand, and are almost infusible. When limestone (shells) is mixed with clay and strongly heated in a kiln, the hard residue on powdering yields **cement**, which hardens with sand or granite and water into a compact mass (concrete) under the action of water.

**Sand** is an oxide of the element silicon; its chemical name is **silica**. It forms salts known as **silicates**. On fusion with potassium or sodium carbonate it drives out the carbon dioxide and forms **water glass** or **soluble**

**glass.** Water glass is used for making cloth and paper fire-proof and also in the manufacture of soap. Eggs are preserved by coating their shells with water glass, which plugs up the pores and makes them air-tight. On decomposing water glass with hydrochloric acid we obtain pure silica. **Glass** is a fused mixture of transparent insoluble silicates.

---

## XXXIV.

### SILVER. GOLD.

**Silver** (Latin, *argentum*) is found in nature combined as well as uncombined. In the uncombined state it occurs often in large crystalline pieces. In the combined state it occurs as sulphide (**silver glance**). Silver sulphide is found in small quantities in **galena**, the chief ore of lead, which is also a sulphide.

Silver glance (sulphide) is carefully roasted, that is, heated strongly in a current of air so that it becomes converted into silver sulphate. This is dissolved in water, and from the solution metallic silver is precipitated by the action of copper.

Silver is the whitest and brightest of all metals, softer than copper but harder than gold. It is very malleable and ductile. It melts at  $950^{\circ}\text{C}$ . and its specific gravity is 10.47. It is the best conductor of heat and electricity.

In the air it remains unchanged. In contact with sulphur compounds silver turns black owing to the formation of the sulphide.

Silver dissolves easily in nitric acid with evolution of reddish-brown fumes. On evaporating the solution colourless crystals of **silver nitrate** are obtained. Silver nitrate is caustic and is employed in medicine for cauterising the flesh. It blackens all vegetable and animal substances. A drop of a solution of silver nitrate falling on the skin produces a black stain.

Silver chloride, bromide and iodide are sensitive to light and therefore find employment in photography. Dry plates contain silver bromide and gelatine.

Silver currency contains copper. White pure silver dissolves in nitric acid to a colourless solution, but a silver coin always gives a blue solution, indicating the presence of copper in it.

**Exp.**—Prove the presence of copper in a two anna piece by the iron and ammonia tests.

Other metals are often coated with silver by dry plating or electro-plating.

**Gold** (Latin, *aurum*), as it has small affinity for other elements, is always found uncombined. Its source is usually either quartz rock, called **auriferous quartz**, or sand. In quartz it is usually found in little round grains. Quartz is a variety of silica; sand consists of small grains of quartz.

Gold-bearing quartz is powdered and treated with mercury to dissolve the pure metal, which is then separated from its amalgam by distillation. Its specific gravity being 19.3, gold can be easily separated by elutriation from sand, whose specific gravity rarely exceeds 2.

Gold is a bright yellow metal, which does not tarnish in contact with air or sulphur. It melts at about 1063.56. It is the most malleable of all metals, and can be hammered into thin leaf about 0.0001mm. thick. On account of its softness it is usually alloyed with copper. Pure gold is said to be 24 carats fine. The British gold coinage contains 11 parts of gold and 1 part of copper. What is its fineness? The percentage of gold in an alloy is sometimes estimated by the colour of the streak on touch-stone.

Gold dissolves in a mixture of hydrochloric and nitric acids (**aqua regia**) forming gold chloride which is used in photography and electro-gilding. It is not attacked by any single acid.

Gold alloyed with silver or copper may be separated by dissolving out the latter by nitric acid.

Silver, copper and brass articles are coated with gold either by **fire-gilding** or by **electro-gilding**.

---

## QUESTIONS.

1. Describe how sand and salt could be separated from a mixture of both.

2. Distinguish clearly between physical changes and chemical changes. Give three examples of each.

2. How can muddy water be rendered clear? What process would you employ to find whether clear water is pure?

3. What is meant by a saturated solution? How will you find out whether a given solution is saturated?

4. What experiments will you perform in order to find out whether a colourless liquid is pure water or not?

6. How can you obtain pure water from sea water? Sketch the apparatus you will employ?

7. A white powder, thrown into water, disappears partly? How will you proceed to find out whether the given powder is a pure substance or a mixture?

8. How are crystals of common salt obtained from sea water?

9. Explain the terms: filtrate; mother liquor; melting; dissolving; crystallising.

10. 20 c.c. of a saturated solution weighed 21.52 grams, and on evaporation left a solid residue weighing 4.88 grams. Calculate the solubility of the substance.

11. Describe exactly what happens when red oxide of mercury is strongly heated. Which of the changes that take place do you consider to be physical, and which chemical? Give reasons for your answer.

12. When iron is heated strongly black scales form on its surface. How will you find out whether this change is due to combination or decomposition?

13. Describe what you observe (1) when sulphur is heated alone in a dry test-tube and (2) when it is mixed with iron filings and heated.

14. Describe the properties of the substance formed by heating sulphur and iron filings together. Is it a mixture or a compound, and why?

15. Describe fully an experiment to show that a chemical compound has a constant composition.

16. Describe exactly, with sketches of the apparatus you will employ, how you can prepare oxygen from the air.

17. Explain clearly the statement that matter is indestructible. Describe an experiment to illustrate the truth of the statement.

18. How would you show that when phosphorus burns in the air the new substance formed weighs more than the original phosphorus?

19. Give reasons for regarding the air as a mixture of oxygen and nitrogen and not a chemical compound.

20. How would you show that water is not an element but a compound?

21. Why is one portion of the air named active? Describe two experiments in support of your answer.

22. Explain the terms; solvent, solute. Name a solvent for each of the following: common salt; sulphur; resin.

23. What is meant by water of crystallisation? Name some substances which crystallise with water of crystallisation and without it? What is decrepitation?

24. What is meant by a precipitate? Give two examples of substances obtained as precipitates.

25. Describe, with sketches of the apparatus employed, two different methods of depriving air of its oxygen and finding the proportion by volume of the other gas present.



26. Describe exactly what happens when a piece of sulphur is lighted and then placed in a jar of oxygen.

27. What is an oxide? Give two examples. Describe the effect on purple litmus solution of the product of combustion in oxygen of each of the following: iron; magnesium; phosphorus.

28. How would you show that air contains water vapour? How does the presence of aqueous vapour in the air affect its weight? In what kind of weather has a litre of air the greatest weight?

29. Why is a lighted taper extinguished in steam which contains oxygen?

30. What is meant by the terms analysis and synthesis in chemistry? How will you carry out the synthesis of red oxide of mercury and the analysis of water?

31. How is oxygen usually prepared in the laboratory? Sketch the apparatus used. Describe the properties of oxygen.

32. Write what you know about the rusting of iron, and describe experiments to confirm your statements. Describe some methods employed to prevent the rusting of iron.

33. Describe briefly three different methods of obtaining hydrogen from water. Describe experiments for showing its most important properties.

34. What is reduction? How may the oxygen be removed from copper oxide?

35. A combustion tube containing copper wire gauze gains 0.24 gram in weight when heated in a current of oxygen. What weight of water will be formed by passing hydrogen through the heated tube?

36. How would you test an unknown gas to find out if it contained uncombined oxygen? If this were not present, how would you test for the presence of oxygen in the combined state?

37. How would you obtain hydrogen from sulphuric acid, caustic soda solution and steam respectively ?

38. How would you prove that carbon and hydrogen are present in a candle ?

39. Is chalk a mixture or a pure substance ? Give reasons for your answer.

40. How would you prove that chalk contains the elements, carbon, oxygen, and calcium ?

41. Explain the action of plants and animals upon the air.

42. How is ammonia gas prepared ? Sketch the apparatus you would employ for collecting a jar of dry ammonia.

43. What is meant by the hardness of water ? How is it caused, and how can it be removed ?

44. Mention three compounds which contain the element sulphur and explain how you recognise the presence of sulphur in each of them.

45. How would you show that hydrogen is lighter than air and that water is formed when hydrogen burns ?

46. Compare the properties of the gases evolved by the action of hydrochloric acid on zinc, marble, and manganese dioxide respectively.

47. You are given four jars with the information that no jar contains more than one of the following gases : oxygen, hydrogen, nitrogen, carbon dioxide. How would you proceed to identify them ?

48. What two gases are employed for bleaching purposes ? How do they act ?

49. Explain what takes place when the following substances are strongly heated :—red lead ; nitre ; sulphur ; iodine ; sal-ammoniac.

50. Distinguish between the two varieties of phosphorus.

51. How are the following substances made :—baking soda ; bleaching powder ; common salt ; sal-ammoniac ; quicklime ; nitric acid ?

52. Name some carbonates which are industrially important, and describe the preparation of one of them.

53. Compare the properties of the gases produced by the action of concentrated sulphuric acid on each of the following substances: copper turnings; common salt; Chile salt-petre.

54. How would you prove the presence of copper in a silver coin?

55. Compare and contrast aluminium, lead, tin and zinc as regards their physical and chemical properties.

56. Describe three cases in which the same chemical substance exhibits different physical properties. Contrast the forms in one case so as to show the differences in properties.

57. How can carbon monoxide be obtained from carbon dioxide? What would happen if copper oxide were heated in a current of carbon monoxide? Sketch any apparatus you would employ for the purpose.

